CHEMISTRY TOPIC

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Cambridge O Level Chemistry

Your notes

1.1 Solids, Liquids and Gases

Contents

- Kinetic Theory
- States of Matter
- Pressure & Temperature in Gases

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Kinetic Theory

Your notes

Solids, Liquids & Gases

Solids

- Solids have a **fixed** volume and shape and they have a high density
- The atoms **vibrate** in position but can't change location
- The particles are packed very closely together in a fixed and regular pattern

Liquids

- Liquids also have a **fixed** volume but adopt the shape of the container
- They are generally less dense than solids (an exception is water), but much denser than gases
- The particles **move** and **slide** past each other which is why liquids adopt the shape of the container and also why they are able to flow freely

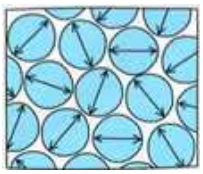
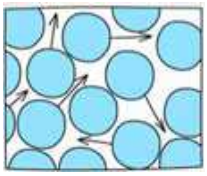
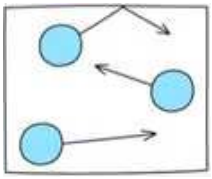
Gases

- Gases do not have a fixed volume and fill their container
- Gases have a very low density
- Since there is a lot of space between the particles, gases can be compressed into a much smaller volume
- The particles are far apart and move randomly and quickly (around 500 m/s) in all directions
- They **collide** with each other and with the sides of the container (this is how **pressure** is created inside a can of gas)

Summary of the Properties of Solids, Liquids and Gases

State	Solid	Liquid	Gas
Density	High	Medium	Low
Arrangement of particles	Regular pattern	Randomly arranged	Randomly arranged
Movement of particles	Vibrate around fixed positions	Move around each other	Move quickly in all directions
Energy of particles	Low energy	Greater energy	Highest energy

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2D diagram			

Your notes

Exam Tip

You can explain the differences in the physical properties of solids, liquids and gases by referring to the arrangement and motion of particles. This is called the kinetic theory of matter.

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States of Matter

Your notes

State Changes

Melting

- Melting is when a solid changes into a liquid
- Requires heat energy which transforms into **kinetic** energy, allowing the particles to move
- Occurs at a specific temperature known as the melting point (m.p.)

Boiling

- Boiling is when a liquid changes into a gas
- Requires heat which causes bubbles of gas to form **below** the surface of a liquid, allowing for liquid particles to escape from the surface and within the liquid
- Occurs at a specific temperature known as the boiling point (b.p.)

Freezing

- Freezing is when a liquid changes into a solid
- This is the reverse of melting and occurs at exactly the **same temperature** as melting, hence the melting point and freezing point of a pure substance are the same. Water, for example, freezes and melts at 0 °C
- Requires a significant decrease in temperature (or loss of thermal energy) and occurs at a specific temperature

Evaporation

- Evaporation occurs when a liquid changes into a gas and occurs over a **range** of temperatures
- Evaporation occurs only at the **surface** of liquids where high energy particles can escape from the liquid's surface at **low** temperatures, below the b.p. of the liquid
- The larger the surface area and the warmer the liquid surface, the more quickly a liquid can evaporate

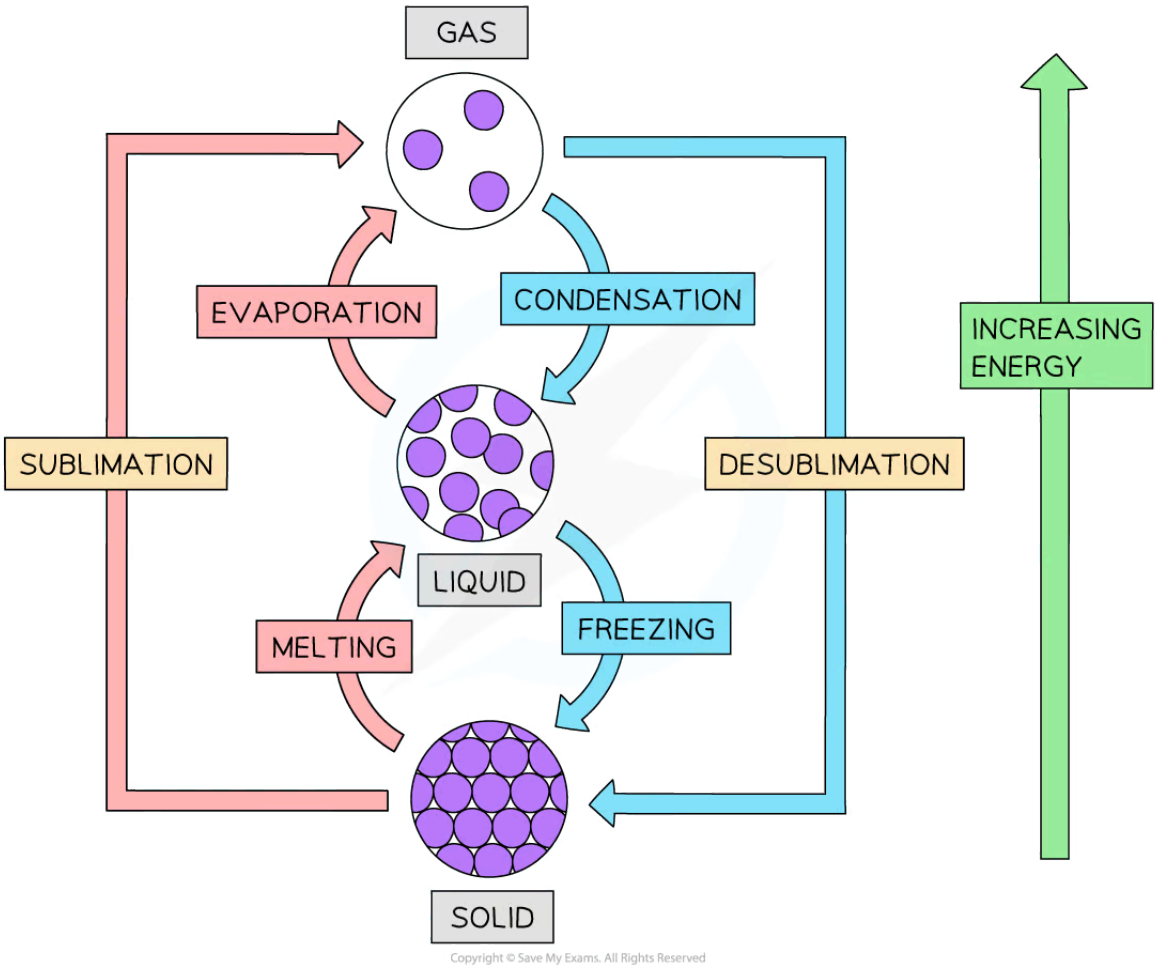
Condensation

- Condensation occurs when a gas changes into a liquid on cooling and it takes place over a **range** of temperatures
- When a gas is cooled its particles lose energy and when they bump into each other they lack the energy to bounce away again, instead they group together to form a liquid

Sublimation

- Sublimation occurs when a solid changes directly into a gas
- This only happens to a few solids, such as iodine or solid carbon dioxide
- The reverse reaction also happens and is called desublimation or deposition

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Interconversion of solids, liquids and gases

Your notes

Exam Tip

Questions on the particle theory of matter show interconversion of states with a reversible arrow: ⇌, which means that the process can go forwards and backwards. Read the question carefully and pick the direction of the change in state that the question refers to.

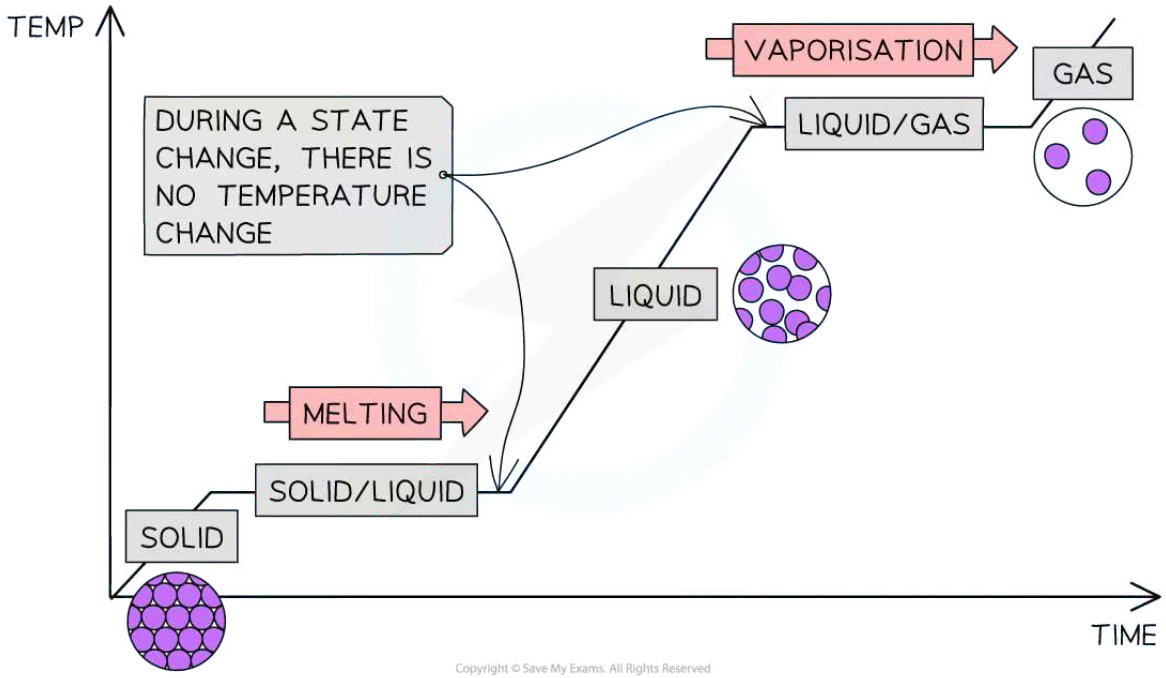
Sublimation is no longer part of the syllabus, but you can see how it fits in on the interconversion diagram with the other state changes.

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State Changes & Kinetic Theory

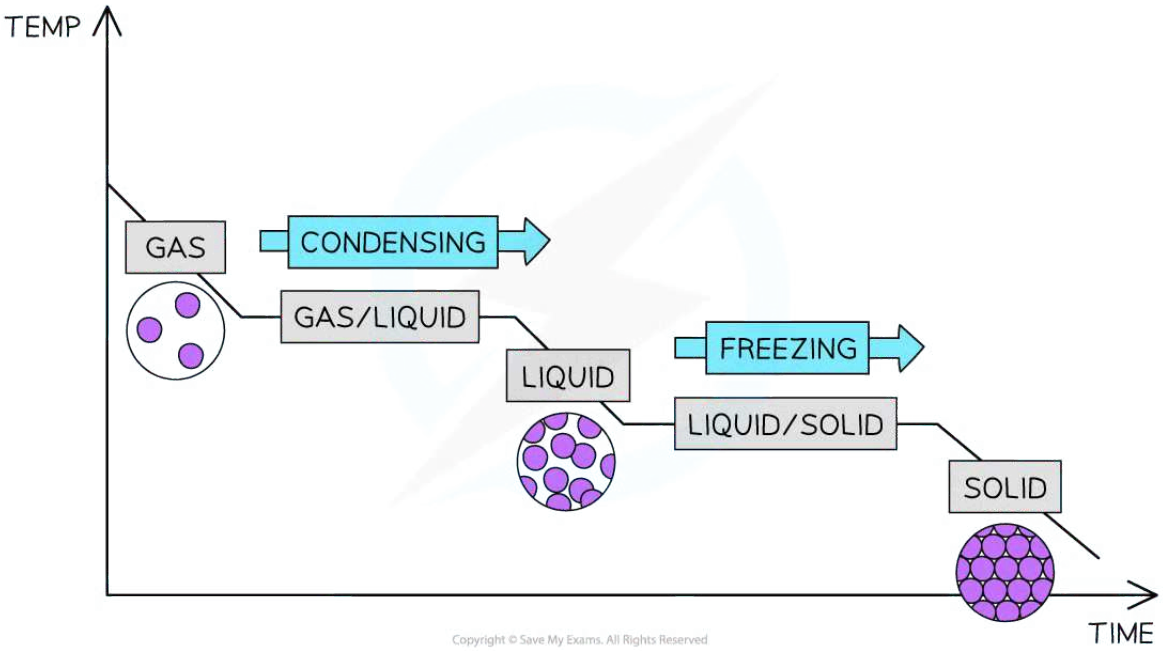
Your notes

- When substances are heated, the particles **absorb** thermal energy which is converted into kinetic energy. This is the basis of the kinetic theory of matter
- Heating a solid causes its particles to vibrate more and as the temperature increases, they vibrate so much that the solid **expands** until the structure breaks and the solid melts
- On further heating, the now liquid substance expands more and some particles at the surface gain sufficient energy to overcome the intermolecular forces and **evaporate**
- When the b.p. temperature is reached, all the particles gain enough energy to escape and the liquids boils
- These changes in state can be shown on a graph called a **heating curve**
- Cooling down a gas has the reverse effect and this would be called a **cooling curve**
- These curves are used to show how changes in temperature affect changes of state
- The horizontal sections occur when there is a change of state but there is no change in temperature



A heating curve showing the states, state changes and temperature changes as time progresses

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A cooling curve is like a heating curve, but is the mirror image

Your notes

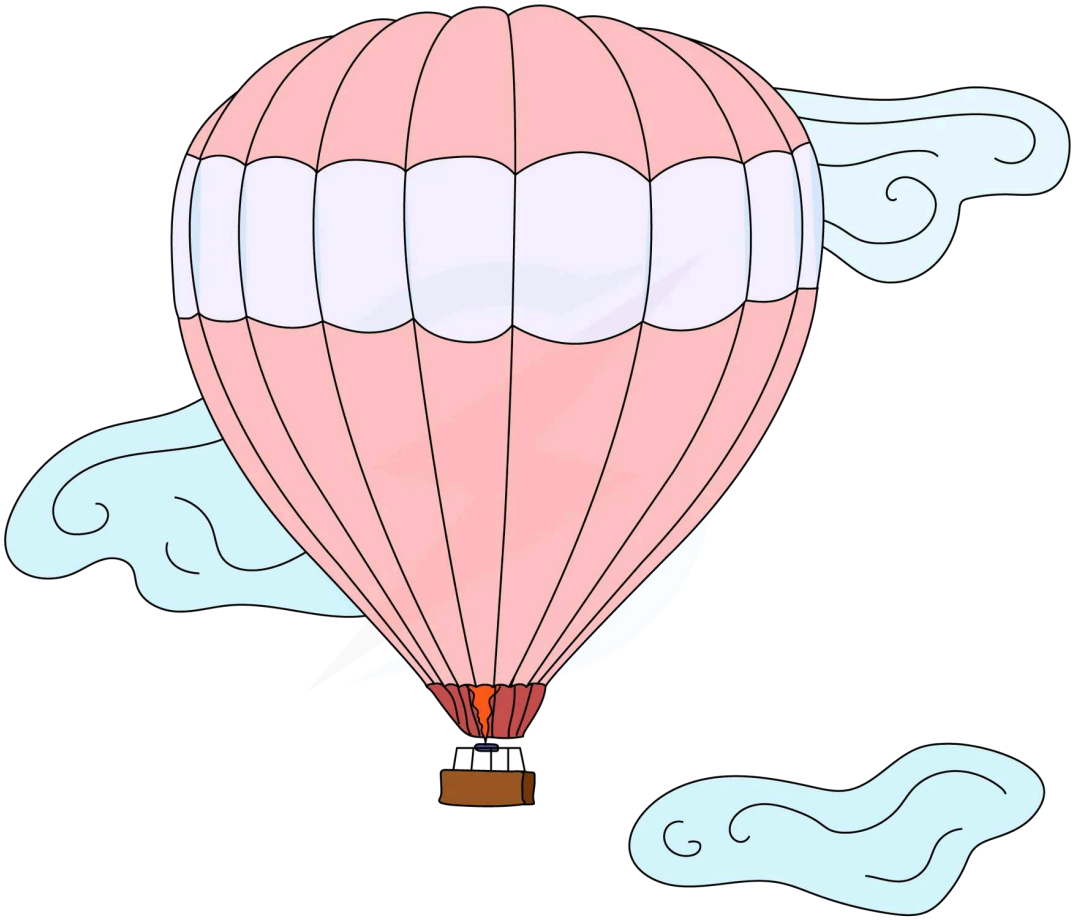
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Pressure & Temperature in Gases

Your notes

Pressure & Temperature in Gases

- A change in temperature or pressure affects the volume of gases
- As the air inside a hot air balloon is heated up, it expands and the balloon gets bigger
- This is because the volume of a gas increases as its temperature increases



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As temperature increases gas volume increases. The density decreases as the volume increases so the balloon rises.

- If you have a gas stored inside a container that is squeezed, the pressure increases as you decrease the volume

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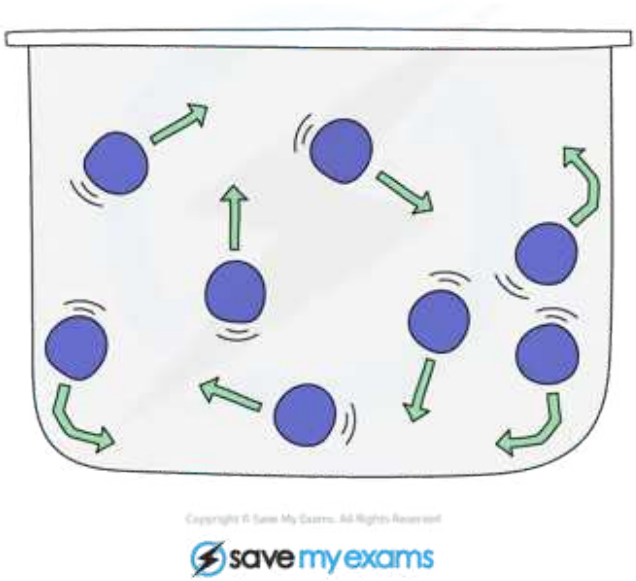
- This is what happens in a bicycle pump
- As you compress the bicycle pump the high pressure allows you to inflate a tire
- You can feel the force of the high pressure if you put your finger on the end of the pump



Pressure increases as volume decreases in a bicycle pump

Kinetic Theory

- Gaseous particles are in constant and **random motion**
- The pressure that gas creates inside a closed container is produced by the gaseous particles hitting the inside walls of the container



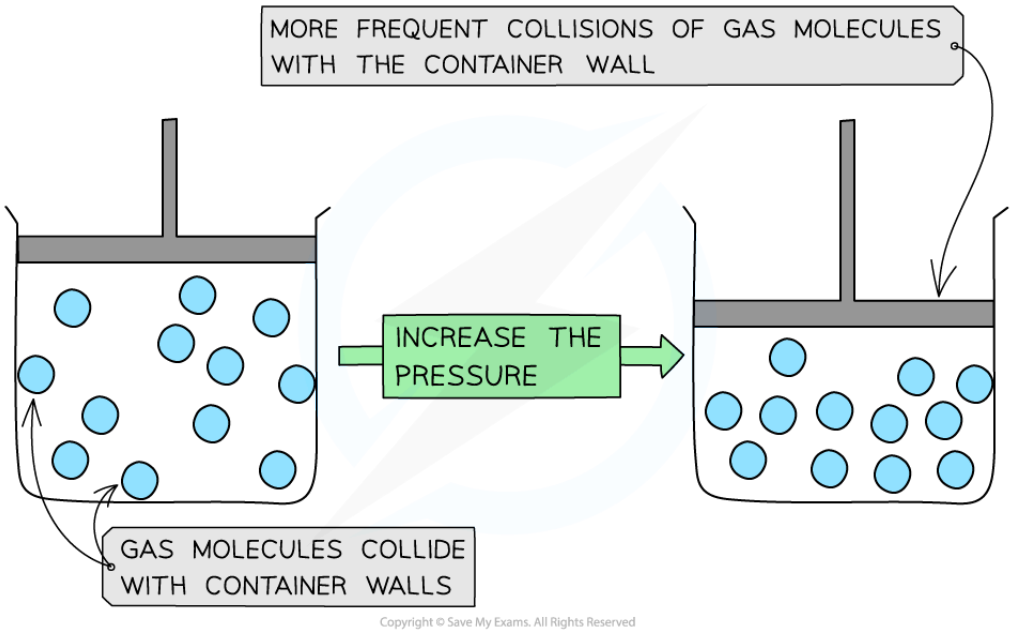
Moving particles of gas colliding with each other and the container walls

Your notes

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- An increase in temperature increases the **kinetic** energy of each particle, as the heat energy is transformed to kinetic energy, so they move faster
- As the temperature increases, the particles in the gas move faster, impacting the container's walls more **frequently**
- If the container walls are flexible and stretchy then the container will get bigger and bigger, just like the hot air balloon!
- If the container is made smaller, then the gas particles hit the wall more frequently
- So when there is a **decrease in volume** this causes an **increase in gas pressure**

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Molecules collide more frequently with the container walls when the pressure is increased

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Cambridge O Level Chemistry

Your notes

1.2 Diffusion

Contents

* Diffusion

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Diffusion

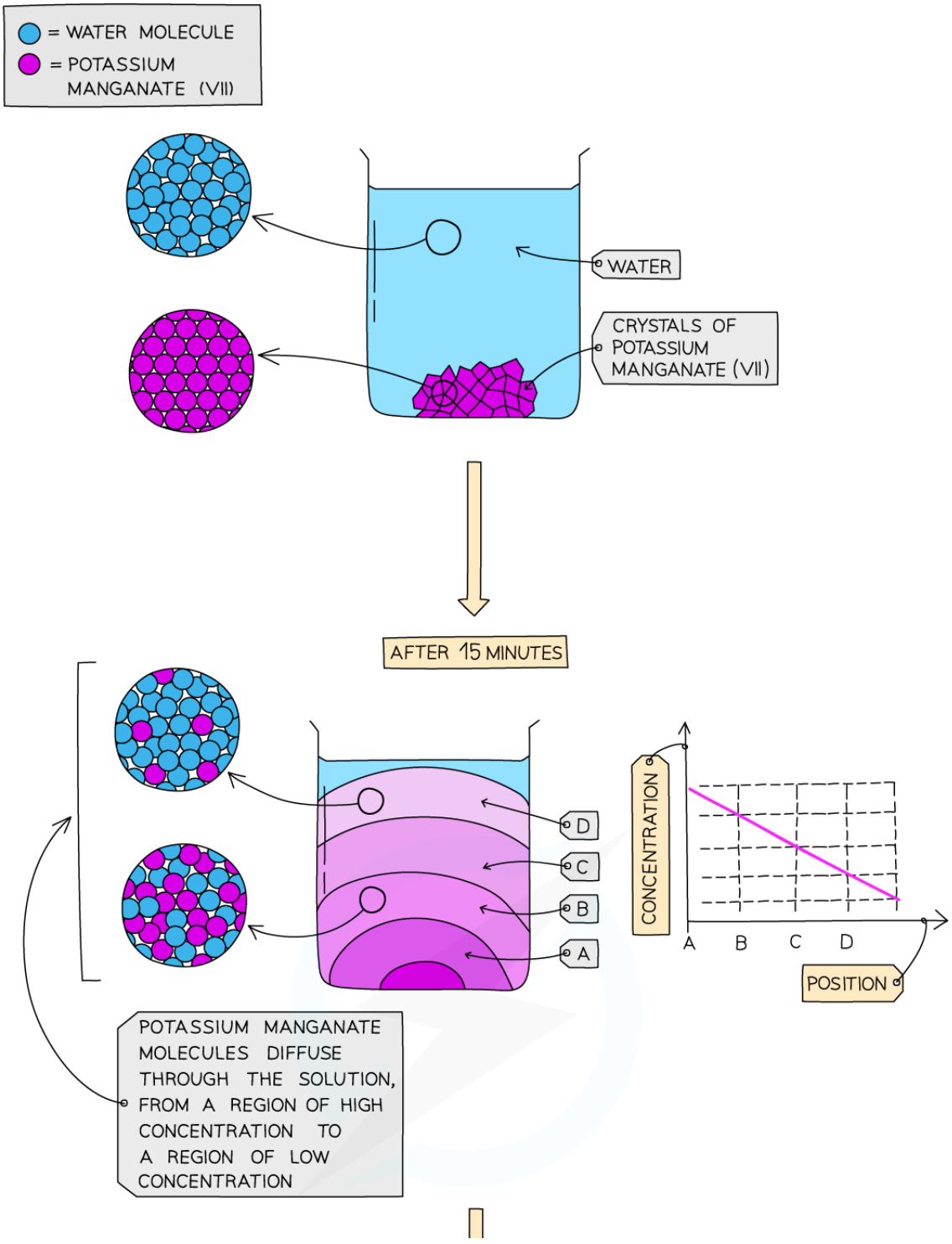

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Diffusion

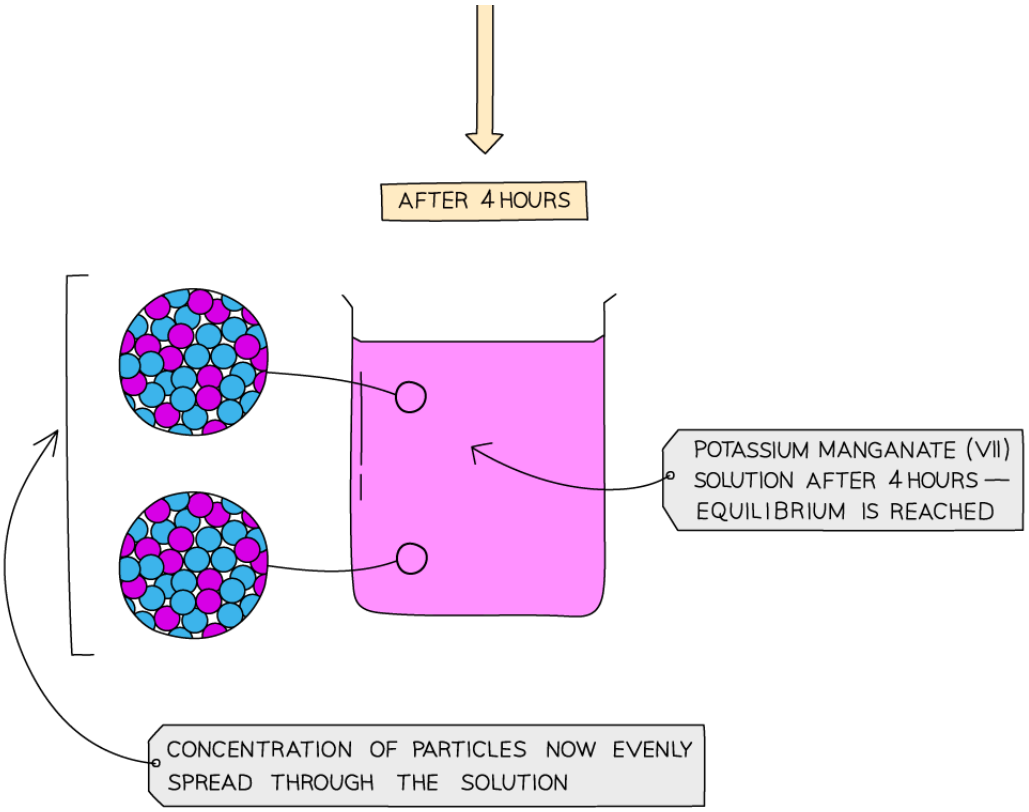
- This is the process by which different gases or different liquids mix and is due to the random motion of their particles
- Diffusing particles move from an area of **high concentration** to an area of **low concentration**
- Eventually the concentration of particles is **even** as they spread out to occupy all of the available space
- Diffusion happens on its own and no energy input is required although it occurs faster at higher temperatures

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Diffusion of potassium manganate(VII), $KMnO_4$, in water. After a few hours the concentration of $KMnO_4$ is the same throughout the solution

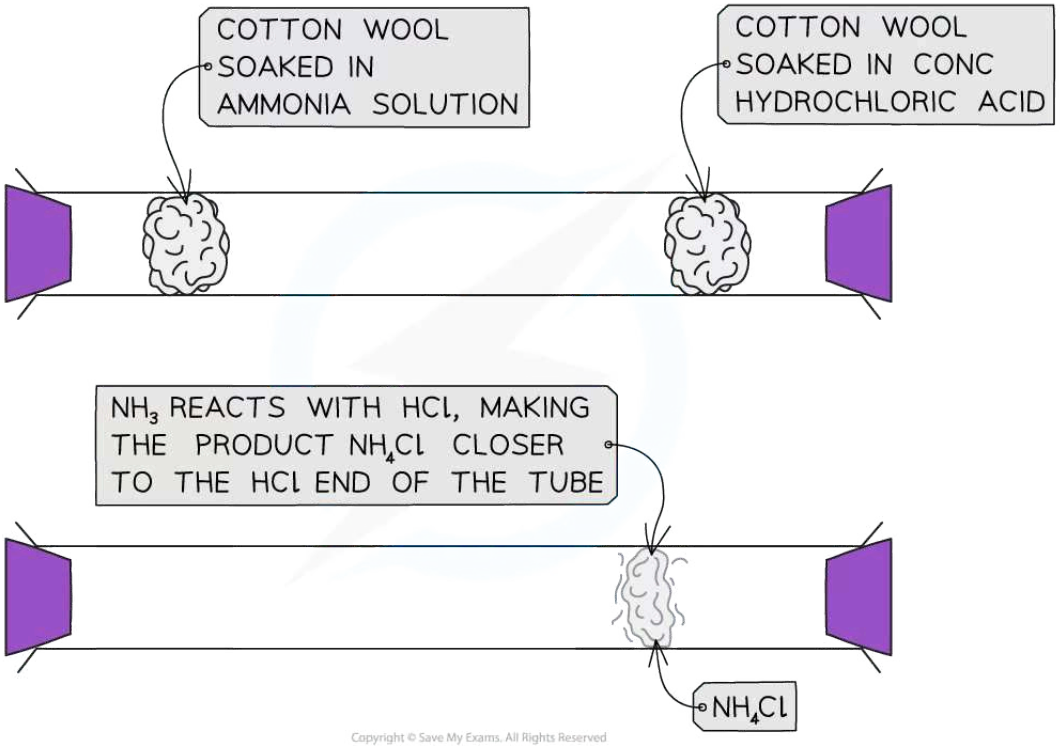
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Diffusion & Molecular Mass

Your notes

- Diffusion occurs much **faster** in gases than in liquids as gaseous particles move much quicker than liquid particles
- At the same temperature, different gases do not diffuse at the same rate.
- This is due to the difference in their relative molecular masses
- Lighter gas particles can travel faster and hence further, therefore the lower its relative mass the faster a gas will diffuse
- This can be demonstrated in the reaction between ammonia, NH_3, and hydrogen chloride gas, HCl, inside a long glass tube
- Where the two gases meet a white smoke of ammonium chloride, NH_4Cl, is formed
- This does not occur in the middle of the tube as you might expect, but much closer to the end with the hydrogen chloride (M_r = 36.5) and the ammonia (M_r = 17) molecules are smaller and lighter



NH_3 molecules have less mass than the HCl molecule, so diffuse faster, hence the product (a white smoke of NH_4Cl) forms closer to the end where the HCl is

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Cambridge O Level Chemistry

Your notes

2.1 Elements, Compounds & Mixtures

Contents

Elements, Compounds & Mixtures

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Elements, Compounds & Mixtures

Your notes

Elements, Compounds & Mixtures

Elements, compounds and mixtures

- All substances can be classified into one of these three types

Element

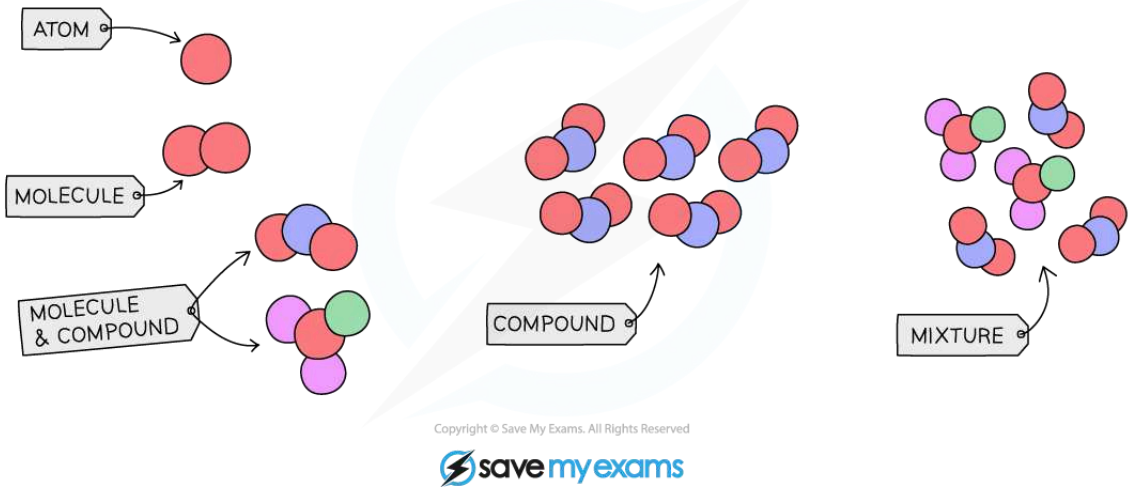
- A substance made of atoms that all contain the **same number of protons** and cannot be split into anything simpler
- There are 118 elements found in the Periodic Table

Compound

- A pure substance made up of two or more elements **chemically combined**
- There is an **unlimited** number of compounds
- Compounds cannot be separated into their elements by physical means
- E.g. copper(II) sulfate ($CuSO_4$), calcium carbonate ($CaCO_3$), carbon dioxide (CO_2)

Mixture

- A combination of two or more substances (elements and/or compounds) that are **not** chemically combined
- Mixtures can be separated by **physical methods** such as filtration or evaporation
- E.g. sand and water, oil and water, sulfur powder and iron filings



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Particle diagram showing elements, compounds and mixtures

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Cambridge O Level Chemistry

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2.2 Atomic Structure & the Periodic Table

Contents

- Atomic Structure
- Electronic Configuration

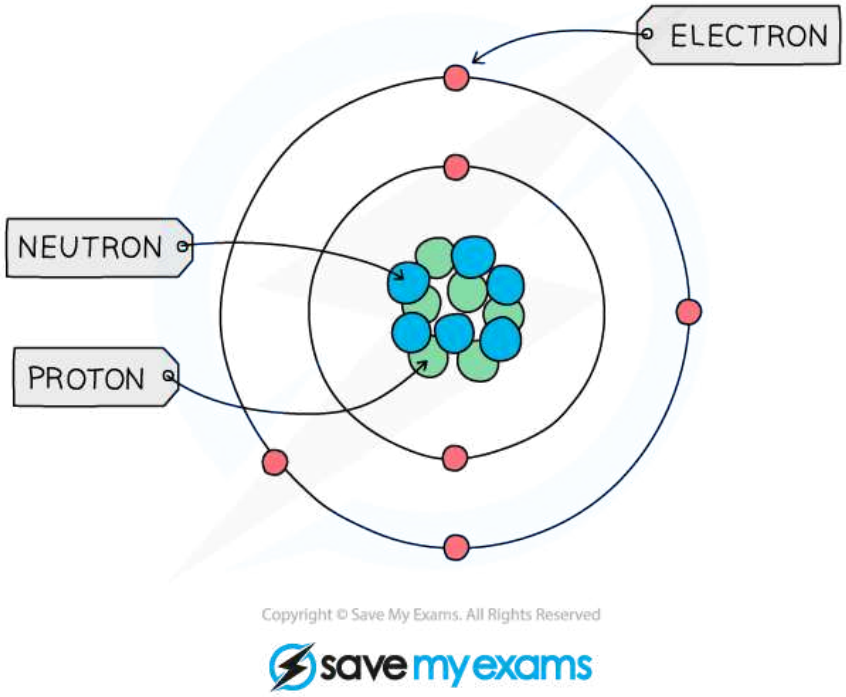
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Atomic Structure

Your notes

Atomic Structure

- All substances are made of tiny particles of matter called **atoms** which are the building blocks of all matter
- Each atom is made of subatomic particles called **protons**, **neutrons,** and **electrons**
- The protons and neutrons are located at the centre of the atom, which is called the **nucleus**
- The electrons move very fast around the nucleus in orbital paths called **shells**
- The mass of the electron is negligible, hence the mass of an atom is contained within the nucleus where the protons and neutrons are located



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The structure of the carbon atom

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Protons, Neutrons & Electrons


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- The size of atoms is so tiny that we can't really compare their masses in conventional units such as kilograms or grams, so a unit called the relative atomic mass is used
- One **relative atomic mass** unit is equal to 1/12th the mass of a carbon-12 atom.
- All other elements are measured relative to the mass of a carbon-12 atom, so relative atomic mass has no units
- Hydrogen for example has a relative atomic mass of 1, meaning that 12 atoms of hydrogen would have exactly the same mass as 1 atom of carbon
- The relative mass and charge of the sub-atomic particles are shown below:

Table of Subatomic Particles

Particle	Relative Mass	Charge
Proton	1	1+
Neutron	1	0 (neutral)
Electron	1/1840	1–

Exam Tip

Knowing the exact mass of an electron is not in the specification and saying it is almost nothing or negligible will be sufficient. It does, however, sometimes appear in particle identification questions, but you can usually deduce that it is the electrons from other information in the question.

Defining Proton Number

- The atomic number (or **proton number**) is the number of protons in the nucleus of an atom
- The symbol for atomic number is ***Z***
- It is also the number of electrons present in a neutral atom and determines the **position** of the element on the Periodic Table

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Defining Mass Number

Your notes

- The Nucleon number (or **mass number**) is the total number of protons **and** neutrons in the nucleus of an atom
- The symbol for nucleon number is ***A***
- The nucleon number **minus** the proton number gives you the number of **neutrons** of an atom
- Note that protons and neutrons can collectively be called **nucleons**.
- The atomic number and mass number of an element can be shown using **atomic notation**
- The **Periodic Table** shows the elements together with their atomic (proton) number at the top and relative atomic mass at the bottom - there is a difference between relative atomic mass and mass number, but for your exam, you can use the relative atomic mass as the mass number (with the exception of chlorine)

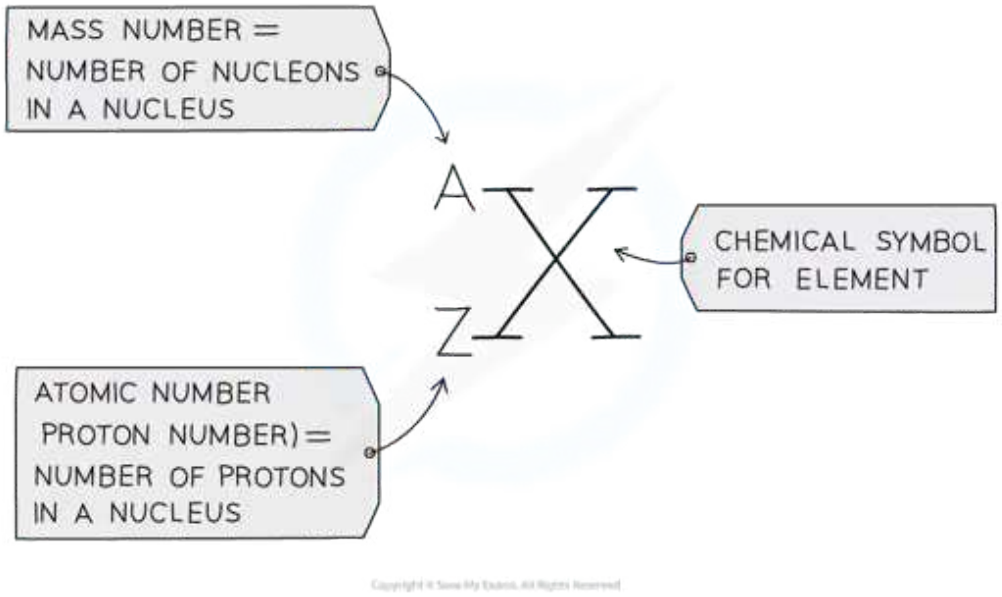
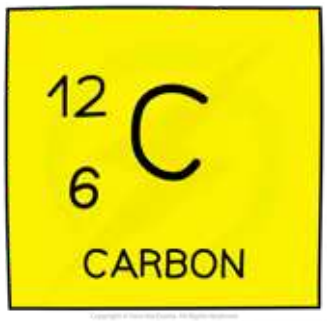


Diagram showing atomic notation

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Your notes



Atomic notation for carbon

Exam Tip

Both the atomic number and the relative atomic number (which you can use as the mass number) are given on the Periodic Table but it can be easy to confuse them. Think **MASS = MASSIVE**, as the mass number is **always** the bigger of the two numbers, the other smaller one is thus the atomic / proton number. **Beware** that some Periodic Tables show the numbers the other way round with the atomic number at the bottom!

Deducing protons, neutrons & electrons

Finding the protons

- The atomic number of an atom and ion determines which element it is
- Therefore, all atoms and ions of the **same element** have the same number of protons (atomic number) in the nucleus
 - E.g. lithium has an atomic number of 3 (three protons) whereas beryllium has atomic number of 4 (four protons)
- The number of protons equals the **atomic (proton) number**
- The number of protons of an **unknown** element can be calculated by using its mass number and number of neutrons:

Mass number = number of protons + number of neutrons

Number of protons = mass number – number of neutrons

Finding the electrons

- An atom is **neutral** and therefore has the **same** number of **protons** and **electrons**

Finding the neutrons

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- The **mass** and **atomic numbers** can be used to find the number of **neutrons** in **ions** and **atoms:**

Number of neutrons = mass number – number of protons


Your notes

Worked example

Determine the number of protons, electrons and neutrons in an atom of element X with atomic number 29 and mass number 63.

Answer:

- The number of protons of element **X** is the same as the atomic number

Number of protons = 29

- The **neutral atom** of element **X** therefore also has **29 electrons**
- The atomic number of an element **X** atom is 29 and its mass number is 63

Number of neutrons = mass number – number of protons

Number of neutrons = 63 – 29

Number of neutrons = 34

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Electronic Configuration

Your notes

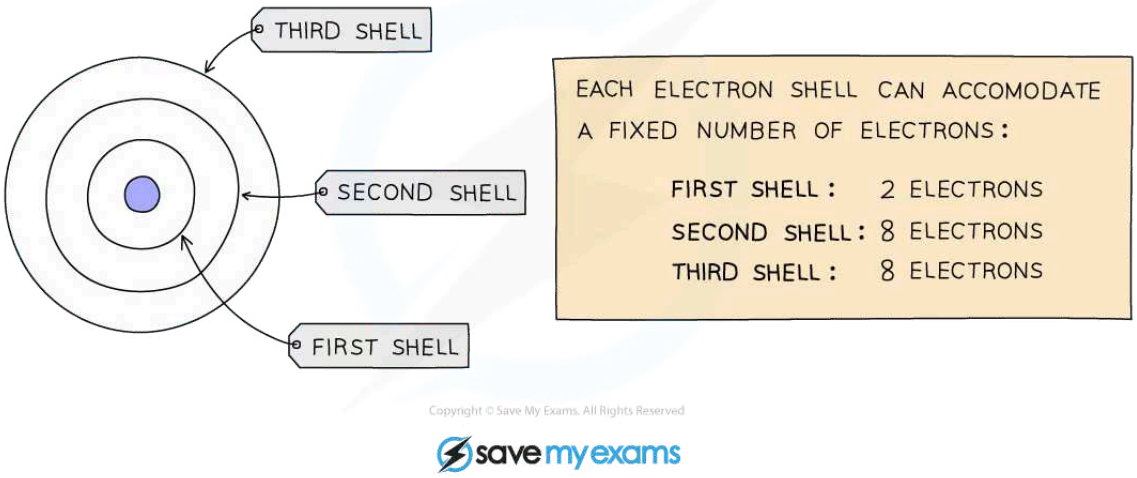
Electronic Configuration

Electronic configuration

- We can represent the structure of the atom in two ways: using diagrams called **electron shell diagrams** or by writing out a special notation called the **electronic configuration** (or electronic structure or electron distribution)

Electron shell diagrams

- Electrons orbit the nucleus in **shells** (or **energy levels**) and each shell has a different amount of energy associated with it
- The further away from the nucleus, the more energy a shell has
- Electrons fill the shell closest to the nucleus
- When a shell becomes full of electrons, additional electrons have to be added to the next shell
- The first shell can hold 2 electrons
- The second shell can hold 8 electrons
- For this course, a simplified model is used that suggests that the third shell can hold 8 electrons
 - For the first 20 elements, once the third shell has 8 electrons, the fourth shell begins to fill
- The outermost shell of an atom is called the **valence** shell and an atom is much more stable if it can manage to completely fill this shell with electrons



A simplified model showing the electron shells

- The arrangement of electrons in shells can also be explained using numbers

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- Instead of drawing electron shell diagrams, the number of electrons in each electron shell can be written down, separated by commas
- This notation is called the electronic configuration (or electronic structure)
 - E.g. Carbon has 6 electrons, 2 in the 1st shell and 4 in the 2nd shell
 - Its electronic configuration is 2,4
- Electronic configurations can also be written for ions
 - E.g. A sodium atom has 11 electrons, a sodium ion has lost one electron, therefore has 10 electrons; 2 in the first shell and 8 in the 2nd shell
 - Its electronic configuration is 2,8


Your notes

The Electronic Configuration of the First Twenty Elements

Element	Atomic Number	Electronic Configuration
hydrogen	1	1
helium	2	2
lithium	3	2,1
berylium	4	2,2
boron	5	2,3
carbon	6	2,4
nitrogen	7	2,5
oxygen	8	2,6
fluorine	9	2,7
neon	10	2,8
sodium	11	2,8,1
magnesium	12	2,8,2

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aluminium	13	2,8,3
silicon	14	2,8,4
phosphorus	15	2,8,5
sulfur	16	2,8,6
chlorine	17	2,8,7
argon	18	2,8,8
potassium	19	2,8,8,1
calcium	20	2,8,8,2

Note: although the third shell can hold up to 18 electrons, the filling of the shells follows a more complicated pattern after potassium and calcium. For these two elements, the third shell holds 8 and the remaining electrons (for reasons of stability) occupy the fourth shell first before filling the third shell.

Electronic Configuration of Ions

- Ions are formed when an atom loses or gains electrons to become stable
 - **Positively charged ions** are formed when an atom **loses** electrons
 - **Negatively charged ions** are formed when an atom **gains** electrons
 - The size of the charge indicates the **number** of electrons that have been lost or gained
- To find the electronic configuration of an ion:
 - Identify the electronic configuration for the atom
 - Identify whether it has lost or gained electrons from its charge
 - Add or remove electrons depending on the charge of the atom

Your notes

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Worked example

1. Give the electronic configuration for the magnesium ion, Mg^{2+}.

Answer:

- A magnesium atom has 12 electrons so has the electronic configuration 2,8,2
- Magnesium has a 2+ charge which means it has lost two electrons
- These electrons are lost from the outer shell so the electronic configuration is 2.8

2. Give the electronic configuration for the chloride ion, Cl^-

Answer:

- A chlorine atom has 17 electrons so has the electronic configuration 2,8,7
- Chlorine has a 1– charge which means it has gained two electrons
- The electronic configuration is therefore 2,8,8

Exam Tip

You need to be able to write the electronic configuration of the first twenty elements and their ions. You may see electronic configurations using full stops or '+' signs instead of commas. You would not be penalised for using full stops.

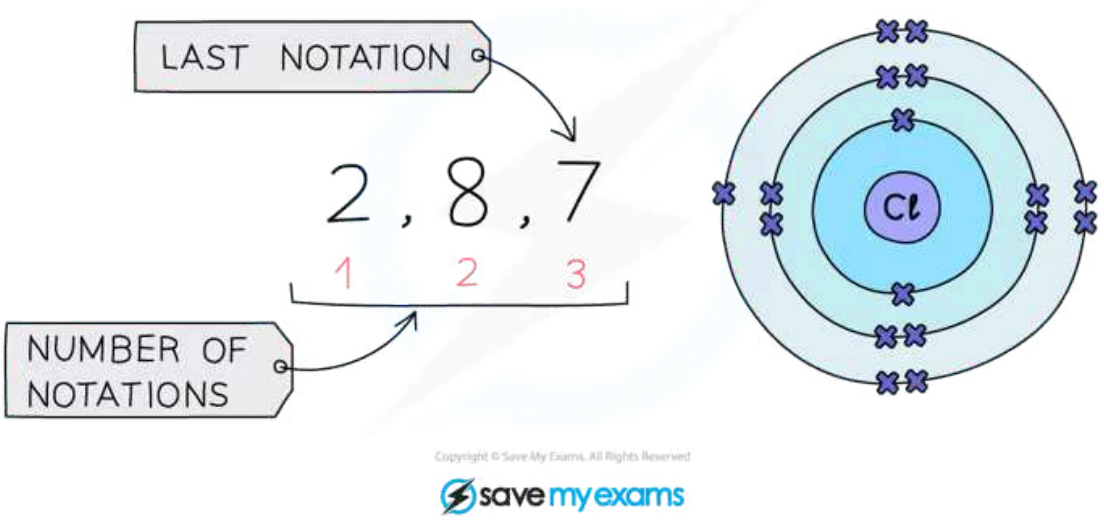
Your notes

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Electron Shells & The Periodic Table

Your notes

- There is a clear **relationship** between the electronic configuration and how the Periodic Table is designed
- The number of notations in the electronic configuration will show the number of occupied shells of electrons the atom has, showing the **period** in which that element is in
- The last notation shows the number of outer electrons the atom has, showing the **group** that element is in (for elements in Groups I to VII)
- Elements in the **same group** have the **same** number of outer **shell electrons**

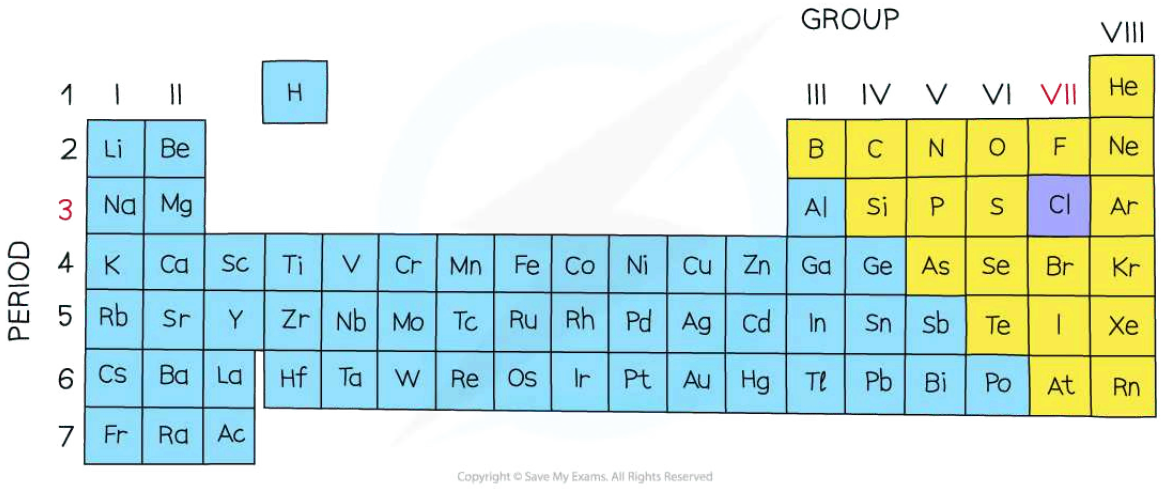


The electronic configuration for chlorine

Period: The red numbers at the bottom show the number of notations which is 3, showing that a chlorine atom has 3 occupied shells of electrons and is in Period 3

Group: The final notation, which is 7 in the example, shows that a chlorine atom has 7 outer electrons and is in Group VII

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The position of chlorine on the Periodic Table

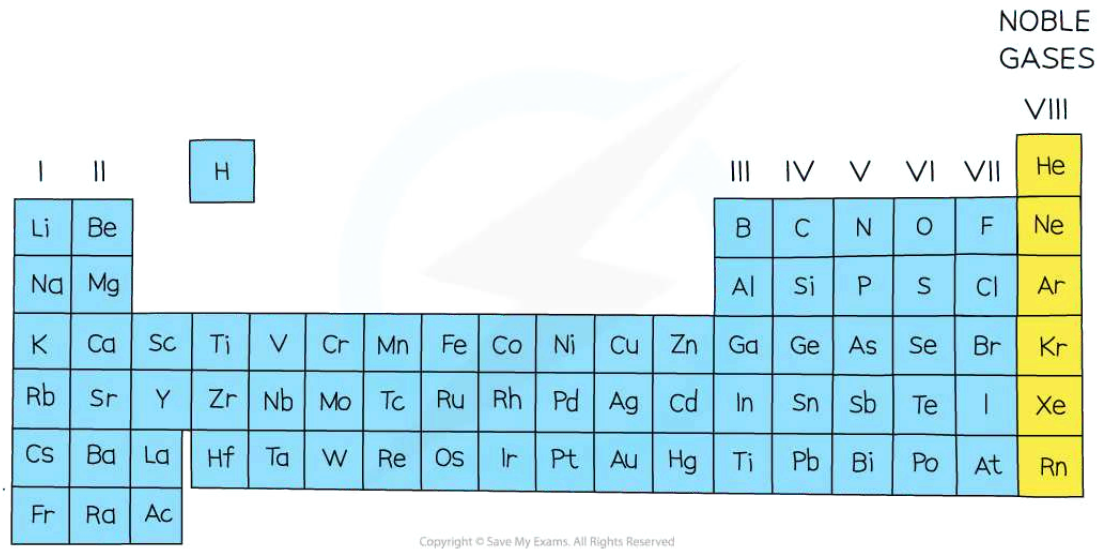
- In most atoms, the outermost shell is not full and therefore these atoms react with other atoms in order to achieve a full outer shell of electrons (which would make them more stable)
- In some cases, atoms lose electrons to entirely empty this shell so that the next shell below becomes a (full) outer shell
- All elements wish to fill their outer shells with electrons as this is a much more stable configuration

The noble gases

- The atoms of the Group VIII elements (the noble gases) all have a full outer shell of electrons
- All of the noble gases are **unreactive** as they have full outer shells and are thus very stable

Your notes

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The noble gases are on the Periodic Table in Group 8/0

Exam Tip

The electrons in the outer shell are also known as valency electrons.

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Cambridge O Level Chemistry

Your notes

2.3 Isotopes

Contents

Isotopes

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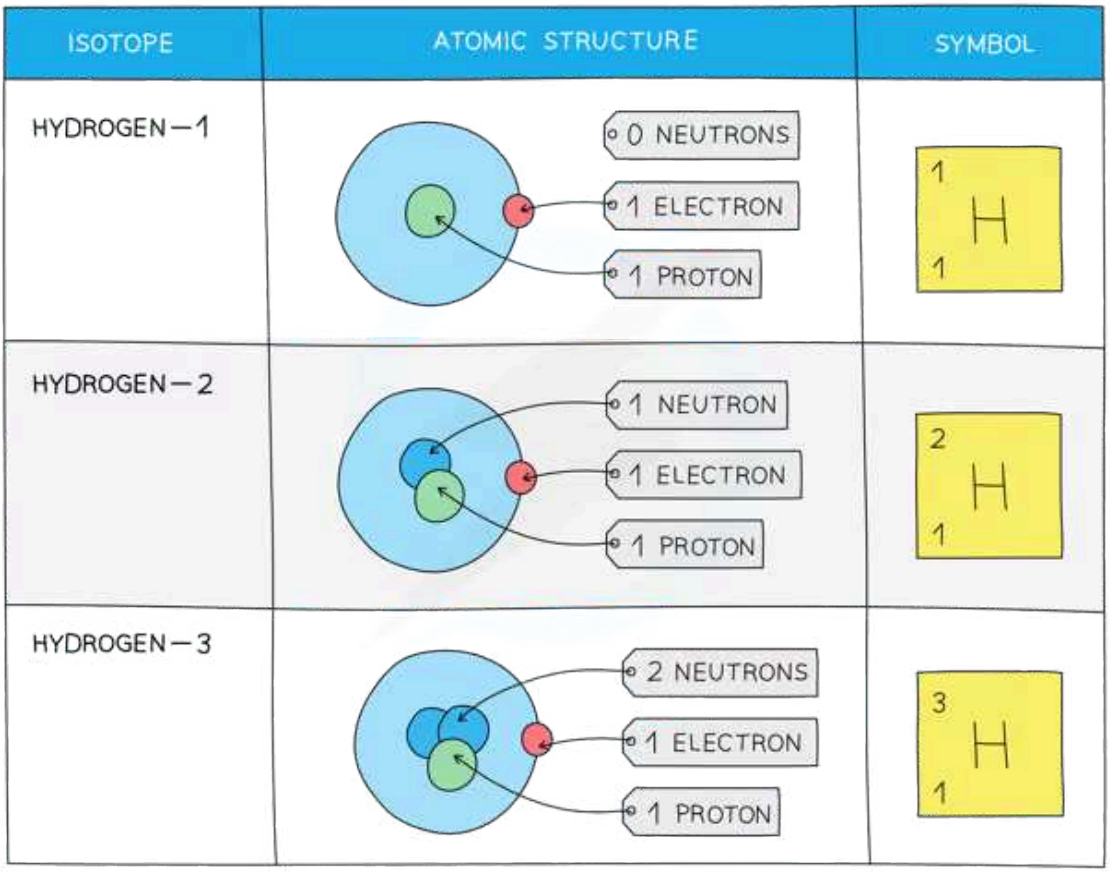
Isotopes

Your notes

Defining Isotopes

- Isotopes are different atoms of the **same element** that contain the same number of **protons** but a different number of **neutrons**
- The symbol for an isotope is the chemical symbol (or word) followed by a dash and then the mass number
- So C-14 (or carbon-14) is the isotope of carbon which contains 6 protons, 6 electrons and 14 - 6 = 8 neutrons
 - It can also be written as ^{14}C or $^{14}_{6}C$

The Atomic Structure and Symbols of the Three Isotopes of Hydrogen



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Why Isotopes Share Properties


Your notes

- Isotopes of the same element display the **same chemical characteristics**
- This is because they have the **same number of electrons in their outer shells** and, therefore, the same electronic configuration and this is what determines an atom's chemistry
- The difference between isotopes is the number of neutrons which are neutral particles within the nucleus and add mass only
- The difference in mass affects the physical properties, such as density, boiling point and melting point
- Isotopes are identical in appearance, so a sample of C-14 would look no different from C-12
- Water made from deuterium oxide is known as 'heavy' water, and has a relative formula of mass 20, compared to 18 for water, so it is 20% heavier, but it would look, taste and feel just like normal water
 - However, it wouldn't be a good idea to drink it because it is toxic as it interferes with biochemical reactions in your cells!

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Calculating Relative Atomic Mass

Relative Atomic Mass

Your notes

- The symbol for the relative atomic mass is A_r
- The relative atomic mass for each element can be found in the Periodic Table along with the atomic number
- The atomic number is shown above the atomic symbol and the relative atomic mass is shown below the atomic symbol
- Atoms are too small to accurately weigh but scientists needed a way to compare the masses of atoms
- The carbon-12 is used as the standard atom and has a fixed mass of 12 units
- It is against this atom which the masses of all other atoms are compared
- Relative atomic mass (A_r) can therefore be defined as:
 - **the average mass of the isotopes of an element compared to 1/12th of the mass of an atom of ^{12}C**
- The relative atomic mass of carbon is 12
 - The relative atomic mass of magnesium is 24 which means that magnesium is twice as heavy as carbon
 - The relative atomic mass of hydrogen is 1 which means it has one-twelfth the mass of one carbon-12 atom
- The relative atomic mass of an element can be calculated from the **mass number** and **relative abundances of all the isotopes** of a particular element using the following equation:

$$A_r = \frac{(\%\text{ of isotope 1} \times \text{mass number of isotope 1}) + (\%\text{ of isotope 2} \times \text{mass number of isotope 2})}{100}$$

- The top line of the equation can be extended to include the number of different isotopes of a particular element present.

Example

The table shows information about the isotopes in a sample of rubidium

ISOTOPE	NUMBER OF PROTONS	NUMBER OF NEUTRONS	PERCENTAGE OF ISOTOPE IN SAMPLE
1	37	48	72
2	37	50	28

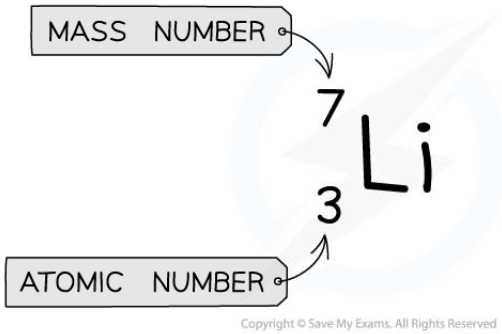
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$$A_r = \frac{(72 \text{ x } 85) + (28 \text{ x } 87)}{100} = 85.6$$

Your notes

Are mass number and relative atomic mass the same thing?

- On the Periodic Table provided in your exam you will see that lithium has a relative atomic mass of **7**
- Although it seems that this is the same as the mass number, they are not the same thing because the relative atomic mass is a **rounded** number
- Relative atomic mass takes into account the existence of isotopes when calculating the mass
- Relative atomic mass is an **average mass** of all the isotopes of that element
- For simplicity **relative atomic masses** are often shown to the nearest whole number



The relative atomic mass of lithium to two decimal places is 6.94 when rounded to the nearest whole number, the RAM is 7, which is the same as the mass number shown on this isotope of lithium

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Cambridge O Level Chemistry

Your notes

2.4 Ions & Ionic Bonds

Contents

- Ions & Ionic Bonds
- Ionic Bonds & Lattice Structure
- Properties of Ionic Compounds

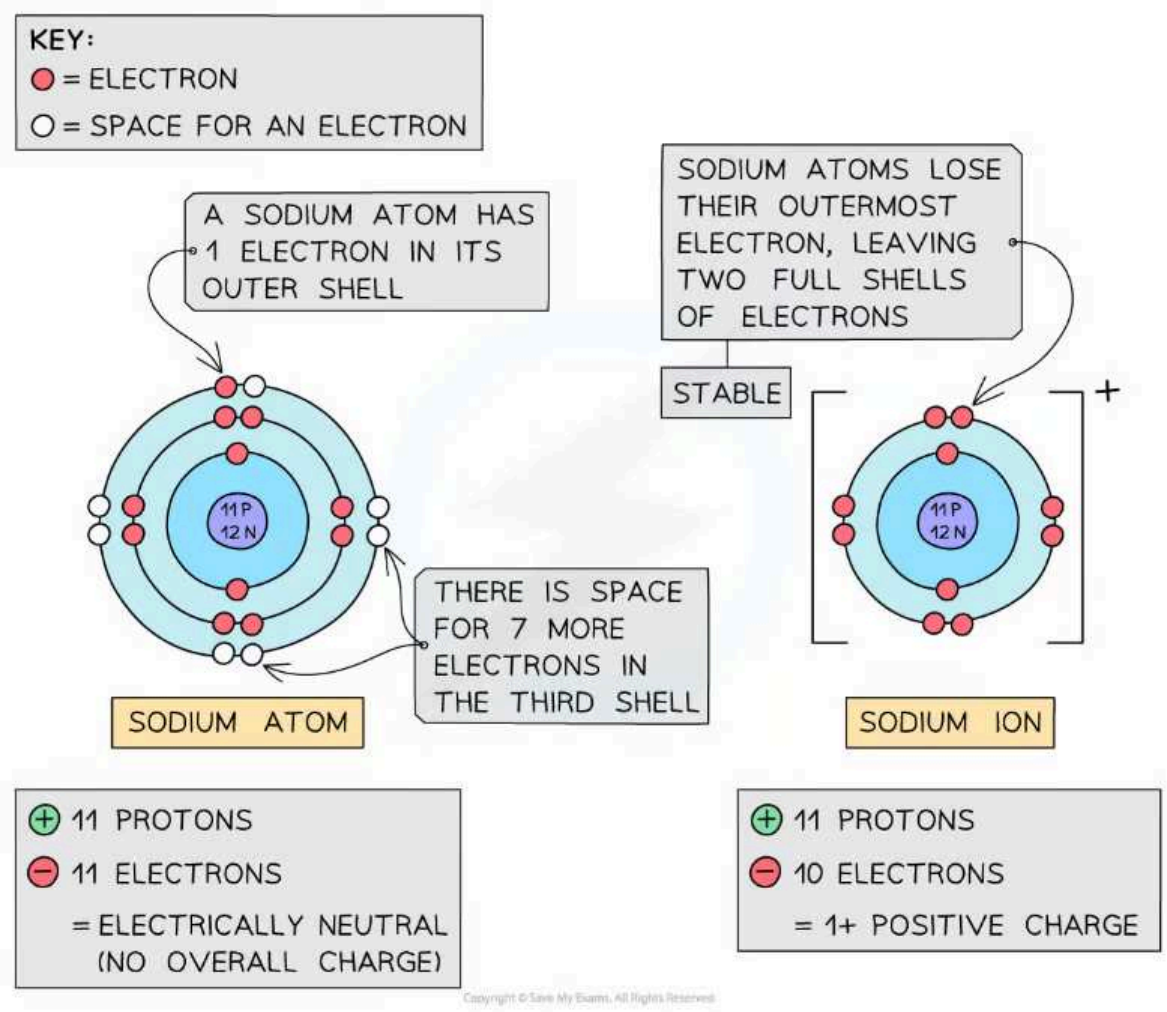
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Ions & Ionic Bonds

Your notes

The Formation of Ions

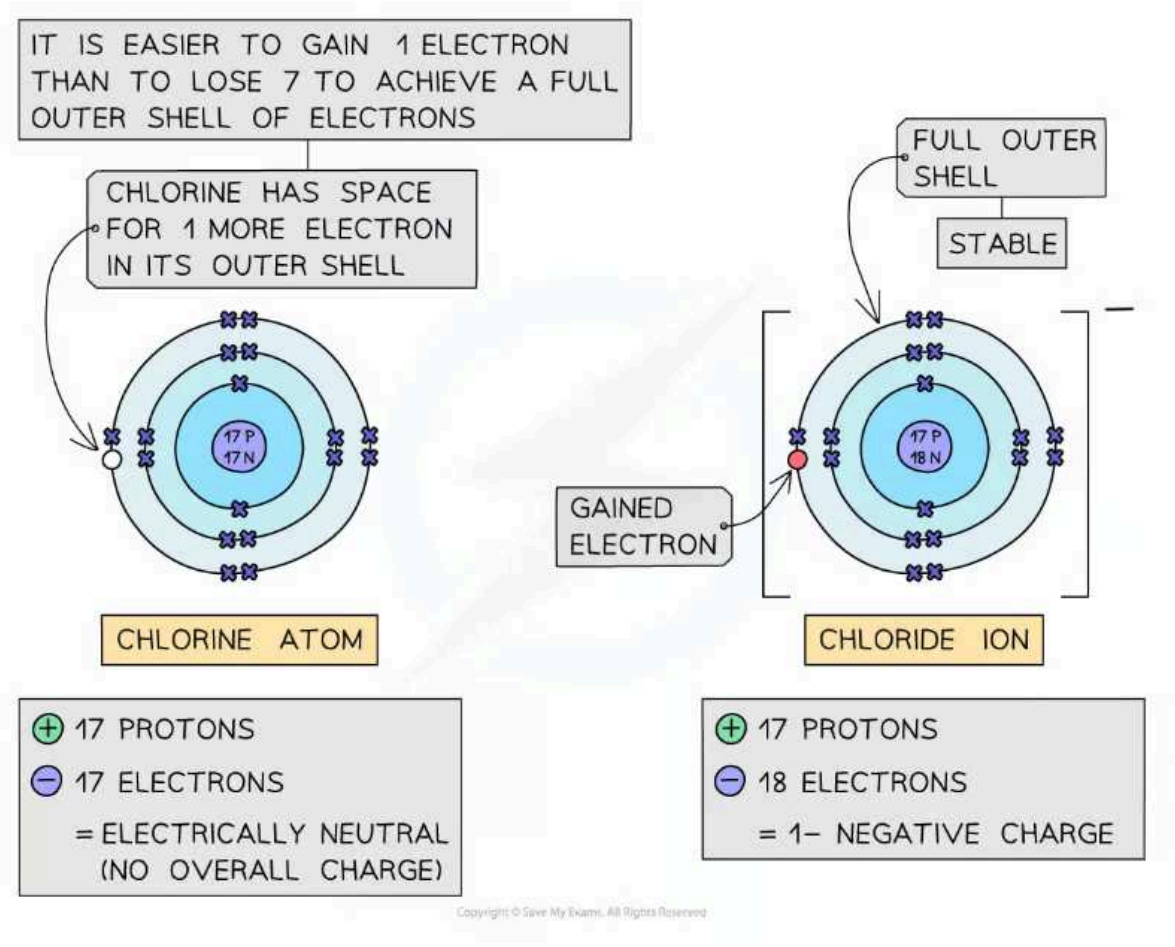
- An ion is an **electrically charged** atom or group of atoms formed by the **loss** or **gain** of **electrons**
- An atom will lose or gain electrons to become more stable
- The loss or gain of electrons takes place to gain a **full outer shell** of electrons which is a more stable arrangement of electrons
- The electronic configuration of an ion will be the same as that of a noble gas – such as helium, neon and argon



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Your notes

Formation of positively charged sodium ion



Formation of negatively charged chloride ion

Ionisation of metals and non-metals

- **Metals:** all metals can **lose** electrons to other atoms to become positively charged ions, known as **cations**
- **Non-metals:** all non-metals can **gain** electrons from other atoms to become negatively charged ions, known as **anions**

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The Formation of Ionic Bonds

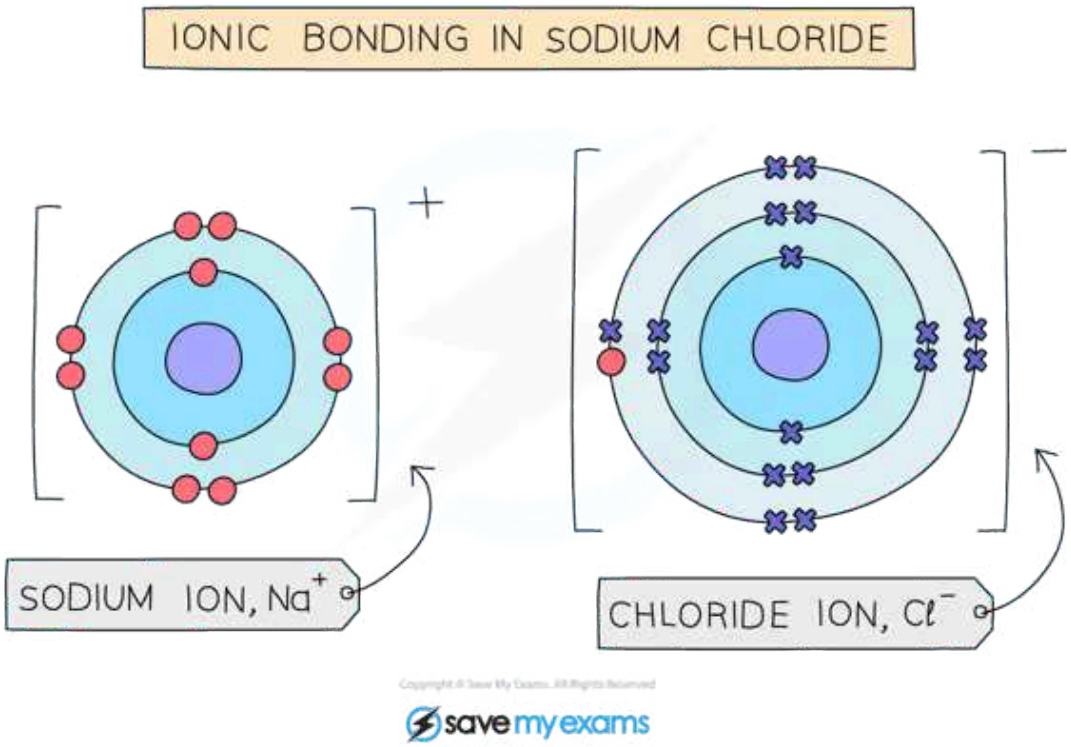
Your notes

- Ionic compounds are formed when metal atoms react with non-metal atoms
- Metal atoms **lose** their outer electrons which the non-metal atoms **gain** to form positive and negative ions
- The positive and negative ions are held together by strong **electrostatic forces of attraction** between **opposite** charges
- This force of attraction is known as an **ionic bond** and they hold ionic compounds together

Dot-and-cross diagrams

- **Dot and cross diagrams** are diagrams that show the arrangement of the outer-shell electrons in an **ionic** or **covalent** compound or element
 - The electrons are shown as dots and crosses

- In a dot and cross diagram:
 - Only the outer electrons are shown
 - The charge of the ion is spread evenly which is shown by using brackets
 - The charge on each ion is written at the top right-hand corner



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Electrostatic forces between the positive Na ion and negative Cl ion

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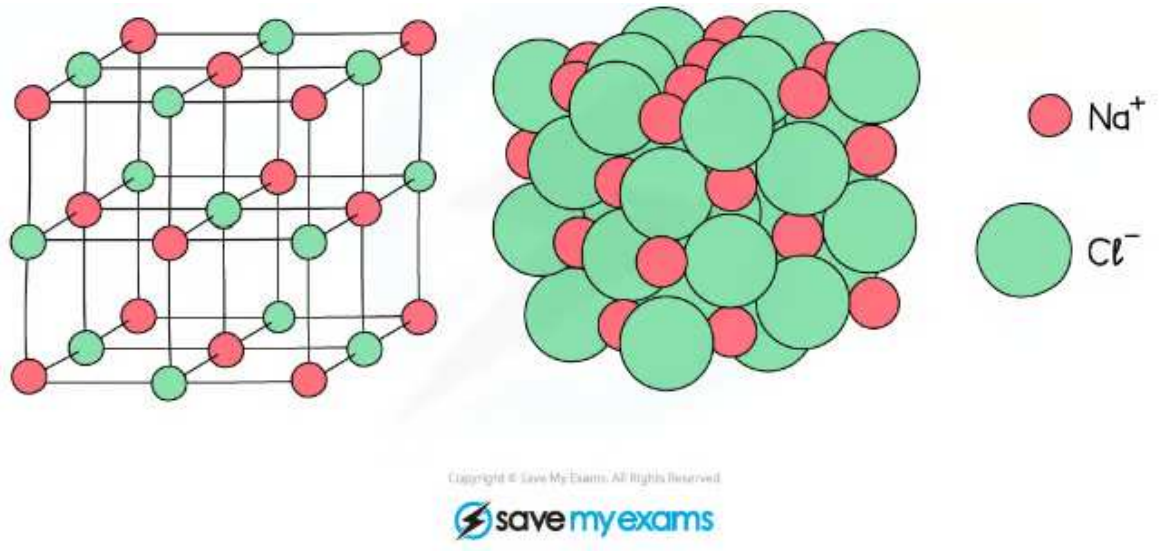
Ionic Bonds & Lattice Structure

Your notes

The Lattice Structure of Ionic Compounds

Lattice structure

- Ionic compounds have a **giant lattice** structure
- Lattice structure refers to the arrangement of the atoms of a substance in 3D space
- In lattice structures, the atoms are arranged in an **ordered** and **repeating** fashion
- The lattices formed by ionic compounds consist of a **regular arrangement** of **alternating** positive and negative ions



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The lattice structure of NaCl

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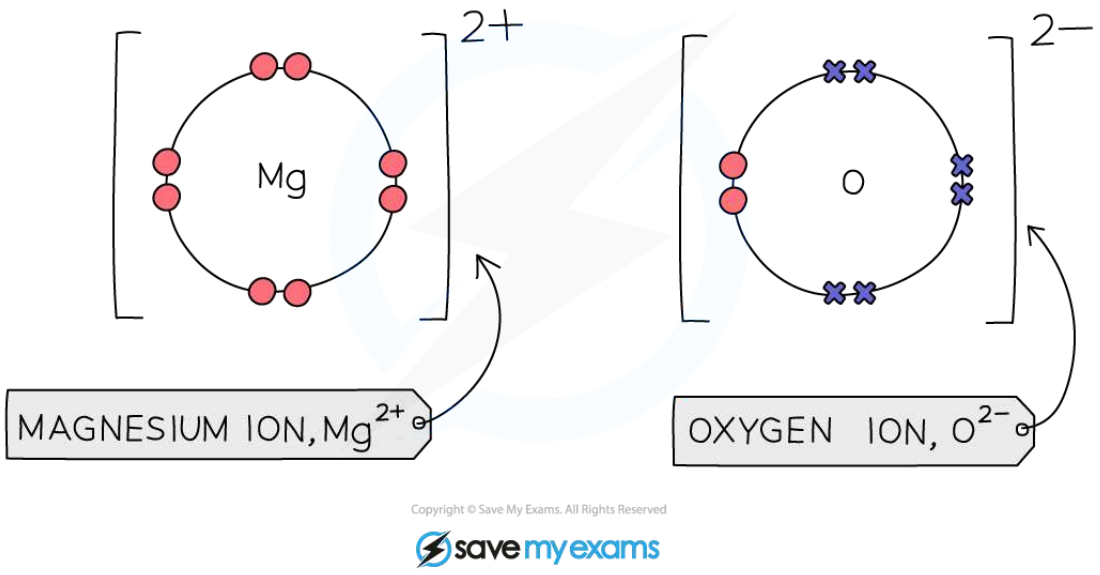
Ionic Bonds between Metallic & Non-Metallic Elements

Your notes

Ionic compounds

- Ionic compounds are formed when metal atoms and non-metal atoms react
- The ionic compound has **no** overall charge

Example: Magnesium Oxide, MgO



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Diagram showing the dot-and-cross diagram of magnesium oxide

Explanation

- Magnesium is a Group II metal so will lose two outer electrons to another atom to have a full outer shell of electrons
- A positive ion with the charge 2+ is formed
- Oxygen is a Group VI non-metal so will need to gain two electrons to have a full outer shell of electrons
- Two electrons will be transferred from the outer shell of the magnesium atom to the outer shell of the oxygen atom
- Oxygen atom will gain two electrons to form a negative ion with charge 2−
- Magnesium oxide has no overall charge

Formula of ionic compound: MgO

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Exam Tip

When drawing dot and cross diagrams, you only need to show the outer shell of electrons. Remember to draw square brackets and include a charge for each ion. Make sure the overall charge is 0; you may need to include more than one positive or negative ion to ensure the positive and negative charges cancel each other out.



Your notes

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Properties of Ionic Compounds


Your notes

Properties of Ionic Compounds

What are the properties of ionic compounds?

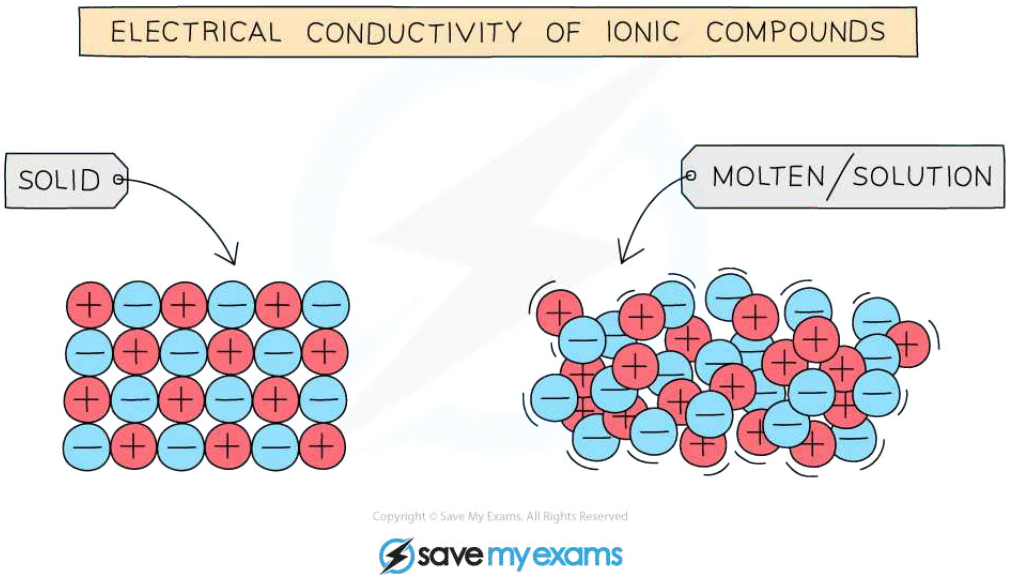
- Ionic compounds are usually **solid** at room temperature
- They have **high** melting and boiling points
- Ionic compounds are good conductors of electricity in the **molten** state or in **solution**
- They are poor conductors in the solid state

Explaining the properties of ionic compounds

- Ionic substances have **high** melting and boiling points due to the presence of **strong electrostatic forces** acting between the oppositely charged ions
- These forces act in all directions and a lot of energy is required to overcome them
- The greater the charge on the ions, the stronger the electrostatic forces and the higher the melting point will be
 - For example, magnesium oxide consists of Mg^{2+} and O^{2-} so will have a higher melting point than sodium chloride which contains the ions, Na^{+} and Cl^{-}
- For electrical current to flow there must be freely moving charged particles such as electrons or ions present
- Ionic compounds are good conductors of electricity in the **molten** state or in **solution** as they have ions that can move and carry a charge
- They are poor conductors in the solid state as the ions are in fixed positions within the lattice and are unable to move

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Cambridge O Level Chemistry

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2.5 Simple Molecules & Covalent Bonds

Contents

- Covalent Bonds
- Properties of Simple Molecular Compounds

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Covalent Bonds


Your notes

The Formation of Covalent Bonds

Covalent compounds

- Covalent compounds are formed when **pairs** of electrons are **shared** between atoms
- Only **non-metal elements** participate in covalent bonding
- As in ionic bonding, each atom gains a **full outer shell** of electrons, giving them a noble gas electronic configuration
- When two or more atoms are **covalently** bonded together, we describe them as '**molecules**'
- **Dot-and-cross** diagrams can be used to show the electric configurations in simple molecules
- Electrons from one atom are represented by a dot, and the electrons of the other atom are represented by a cross
- The electron shells of each atom in the molecule overlap and the shared electrons are shown in the area of overlap
- The dot-and-cross diagram of the molecule shows clearly which atom each electron originated from

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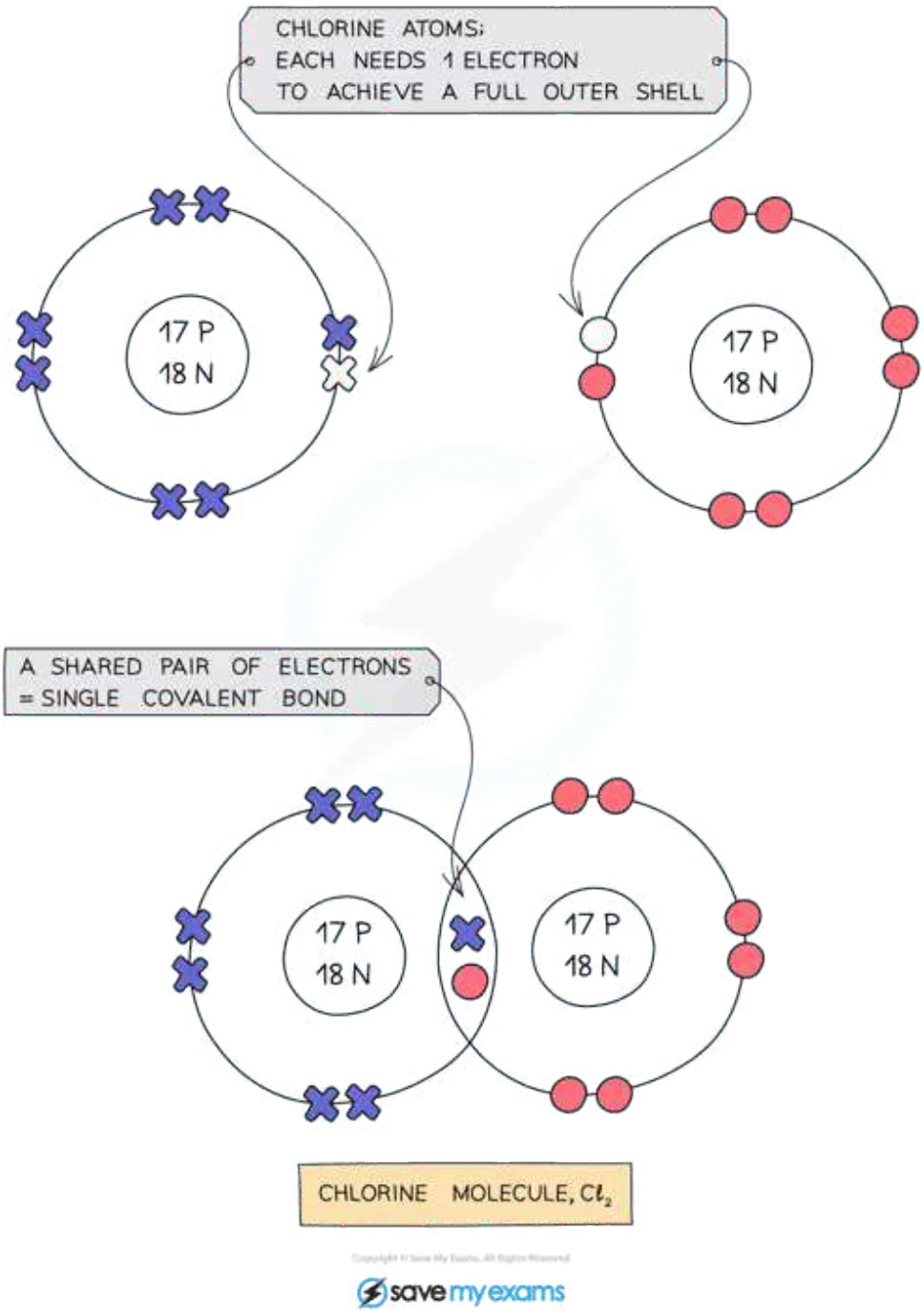


Diagram showing how a covalent bond forms between two chlorine atoms

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Exam Tip

When drawing dot-and-cross diagrams for covalent compounds, make sure that the electron shell for each atom is full (remember that the 1st shell can only hold 2 electrons).



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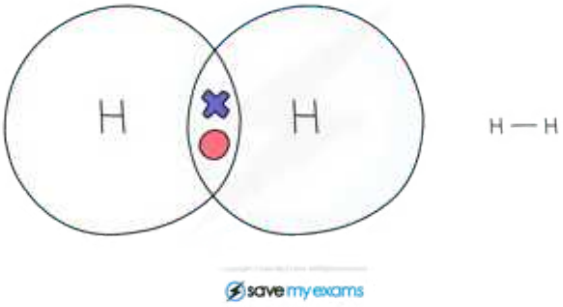
Dot & Cross Diagrams of Simple & Complex Molecules

Your notes

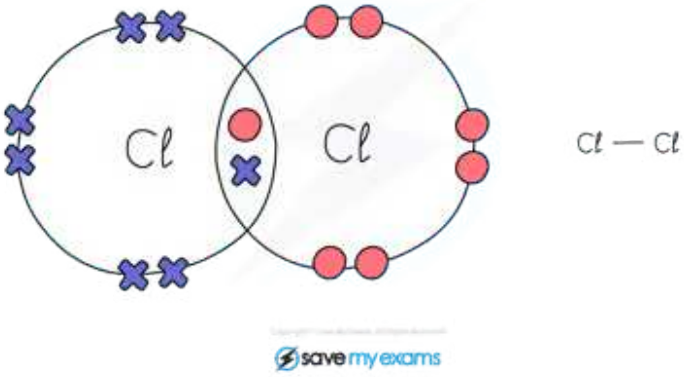
- Many simple molecules exist in which two adjacent atoms share one pair of electrons, also known as a single covalent bond (or single bond)

Common Examples of Simple Molecules

Hydrogen:



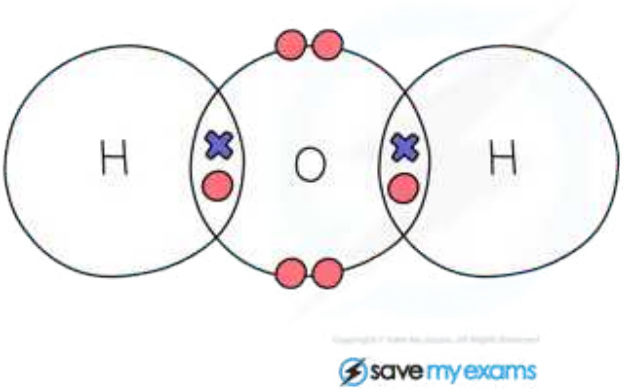
Chlorine:



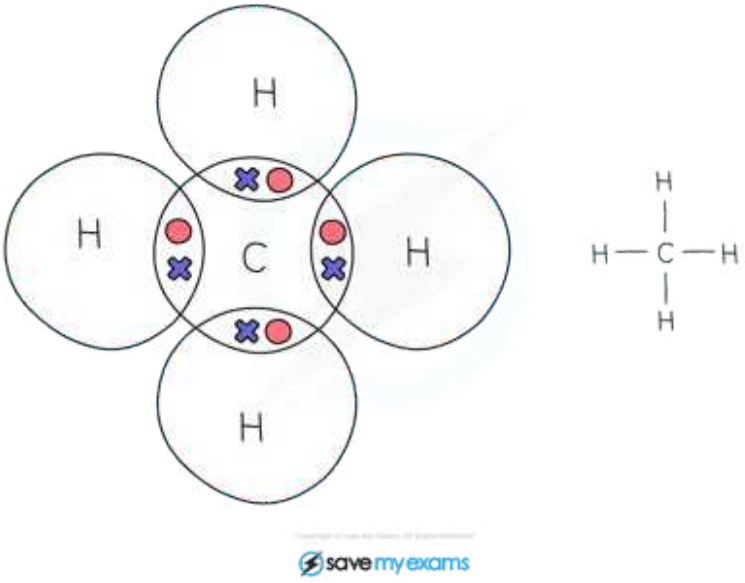
Water:

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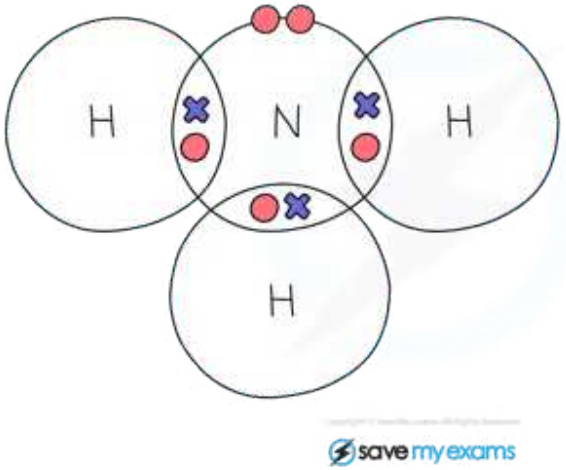
Methane:

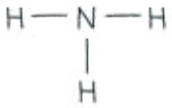


Ammonia:

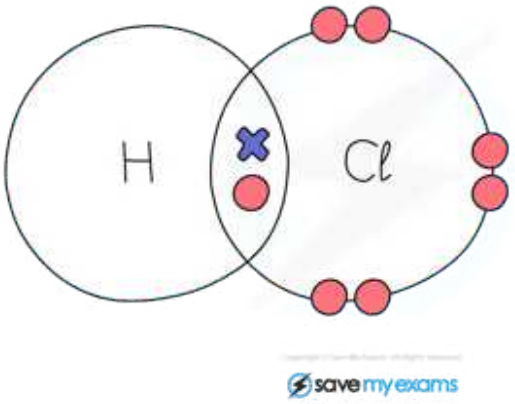
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Hydrogen chloride:





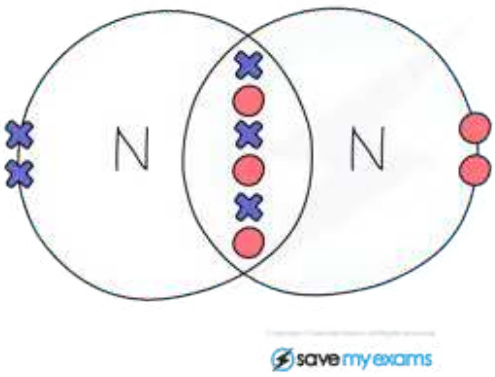
- Some atoms need to share more than one pair of electrons to gain a full outer shell of electrons
- If two adjacent atoms share two pairs of electrons, two covalent bonds are formed, also known as a **double bond**
- If two adjacent atoms share three pairs of electrons, three covalent bonds are formed, also known as a **triple bond**

Nitrogen:

- When 2 nitrogen atoms react they share 3 pairs of electrons to form a triple bond

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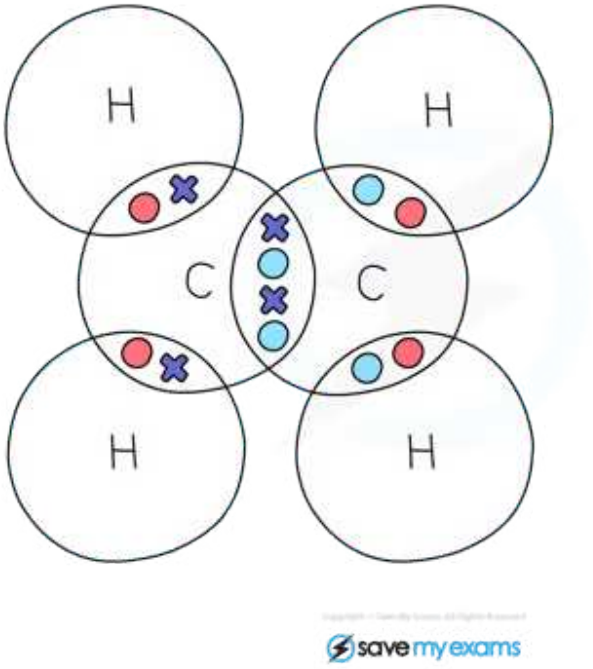
Your notes

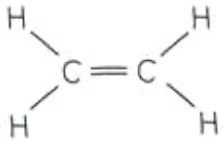


N ≡ N

Ethene:

- In ethene, the 2 carbon atoms share 2 pairs of electrons
- This is known as a double bond

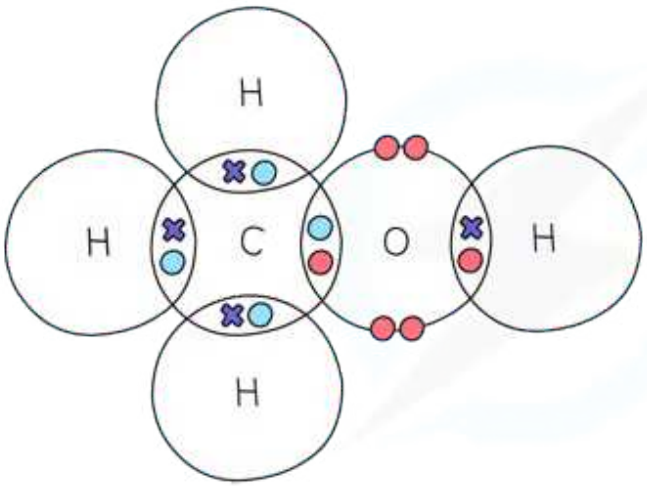


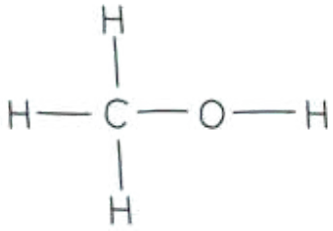


Methanol:

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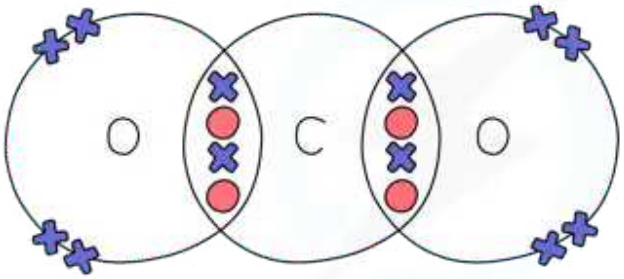
Your notes





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Carbon Dioxide:





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Exam Tip

Be careful when drawing dot-and-cross diagrams, it is a common mistake for students to draw the wrong type of diagram. Remember, if the compound contains **metal** and **non-metal**, it is an **ionic** compound and you need to draw the ions separated, with square brackets around each ion, together with a charge. If the compound contains **non-metal** atoms only, it is a **covalent** compound, the shells should overlap and contain one or more pairs of electrons.

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Properties of Simple Molecular Compounds

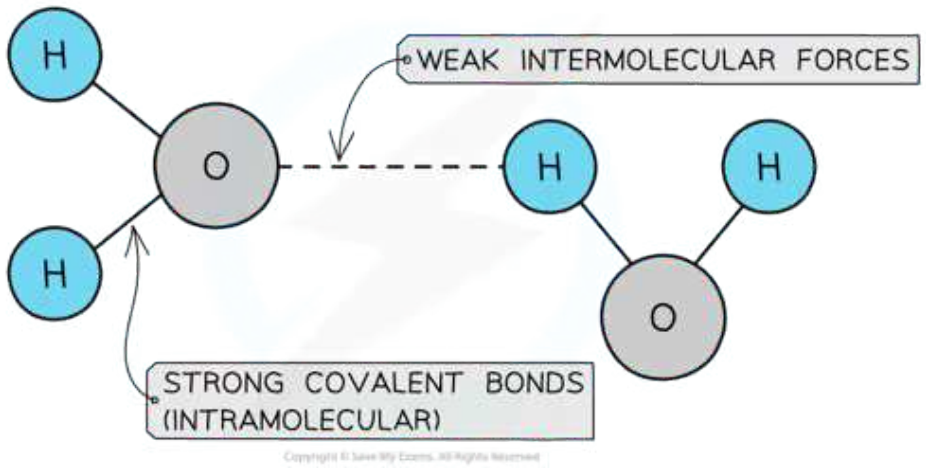
Your notes

Properties of Simple Molecular Compounds

- Small molecules are compounds made up of molecules that contain just a few atoms **covalently bonded** together
- They have **low** melting and boiling points so covalent compounds are usually **liquids** or **gases** at room temperature
- As the molecules increase in **size**, the melting and boiling points generally increase
- Small molecules have poor electrical conductivity

Explaining melting / boiling points

- Small molecules have covalent bonds joining the atoms together, but intermolecular forces that act between neighbouring molecules
- They have **low** melting and boiling points as there are only **weak intermolecular** forces acting **between** the molecules
- These forces are **very weak** when compared to the covalent bonds and so most small molecules are either gases or liquids at room temperature
- As the molecules increase in **size** the intermolecular forces also increase as there are more electrons available
- This causes the melting and boiling points to **increase**



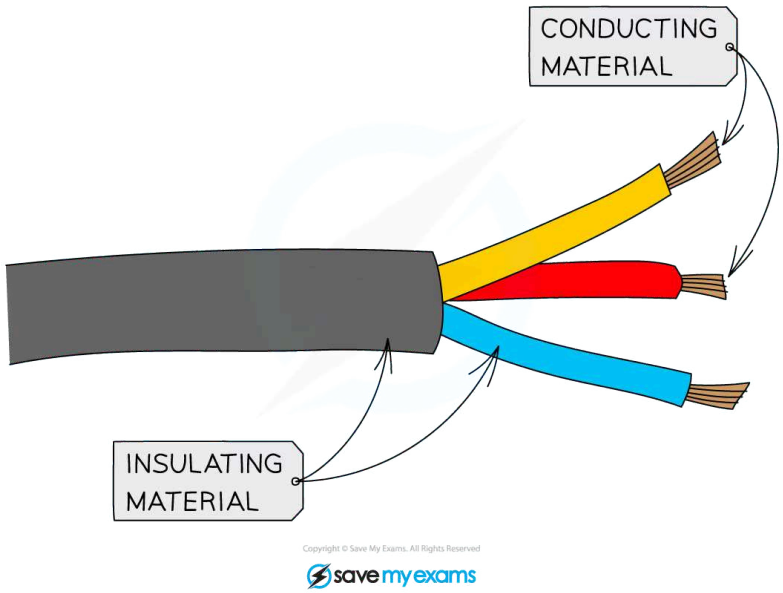
The bonds between hydrogen and oxygen in water are COVALENT, and the attractions between the molecules are INTERMOLECULAR FORCES which are about one tenth as strong as covalent bonds

Explaining electrical conductivity

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Molecular compounds are **poor conductors** of electricity as there are no free ions or electrons to carry the charge.

- Most covalent compounds do not conduct at all in the solid state and are thus **insulators**
- Common insulators include the plastic coating around household electrical wiring, rubber and wood



The plastic coating around electrical wires is made from covalent molecules that do not allow a flow of charge

Exam Tip

The atoms within covalent molecules are held together by covalent bonds while the molecules in a covalent substance are attracted to each other by intermolecular forces.

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2.6 Giant Covalent Structures

Contents

* Diamond, Graphite & Silicon (IV) Oxide

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Diamond, Graphite & Silicon (IV) Oxide

Your notes

Structure of Graphite, Diamond & Silicon(IV) Oxide

- Diamond and graphite are allotropes of carbon which have **giant covalent structures**
- Both substances contain only carbon atoms but due to the differences in bonding arrangements they are physically completely different
- Giant covalent structures contain billions of non-metal atoms, each joined to adjacent atoms by covalent bonds forming a giant lattice structure

Diamond

- In diamond, each carbon atom bonds with four other carbons, forming a **tetrahedron**
- All the covalent bonds are identical, very strong and there are no **intermolecular forces**

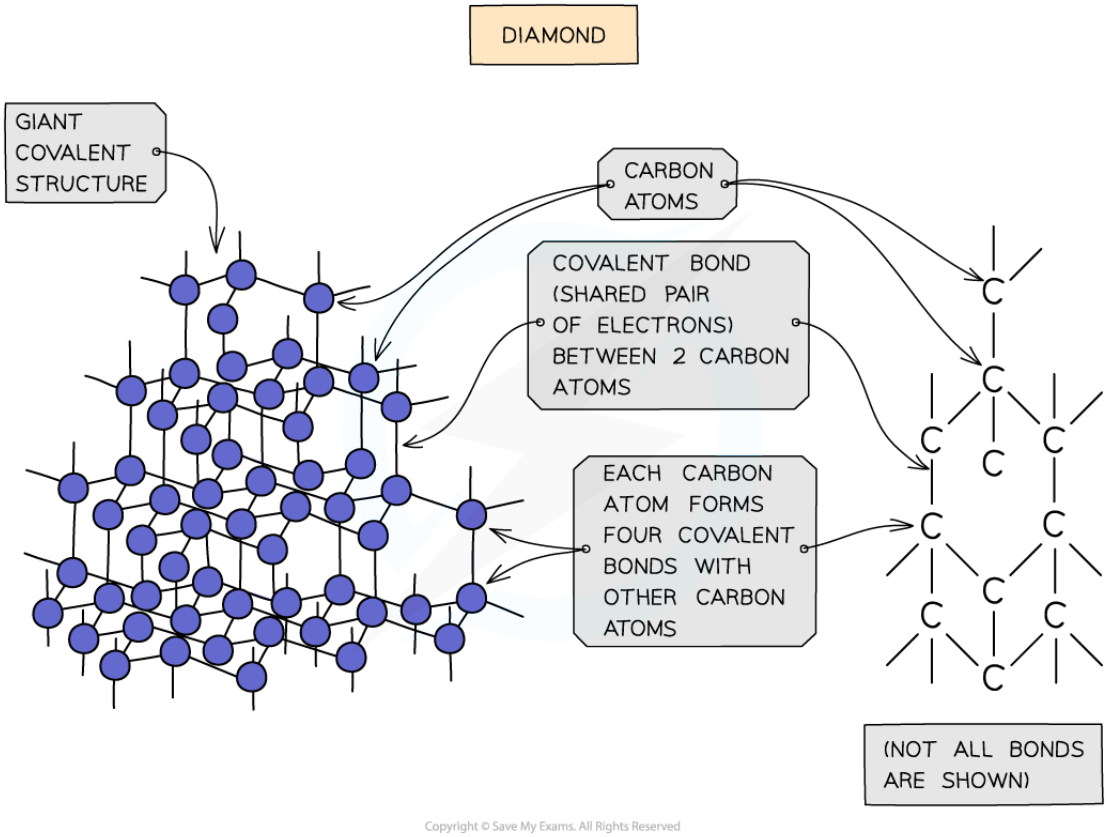


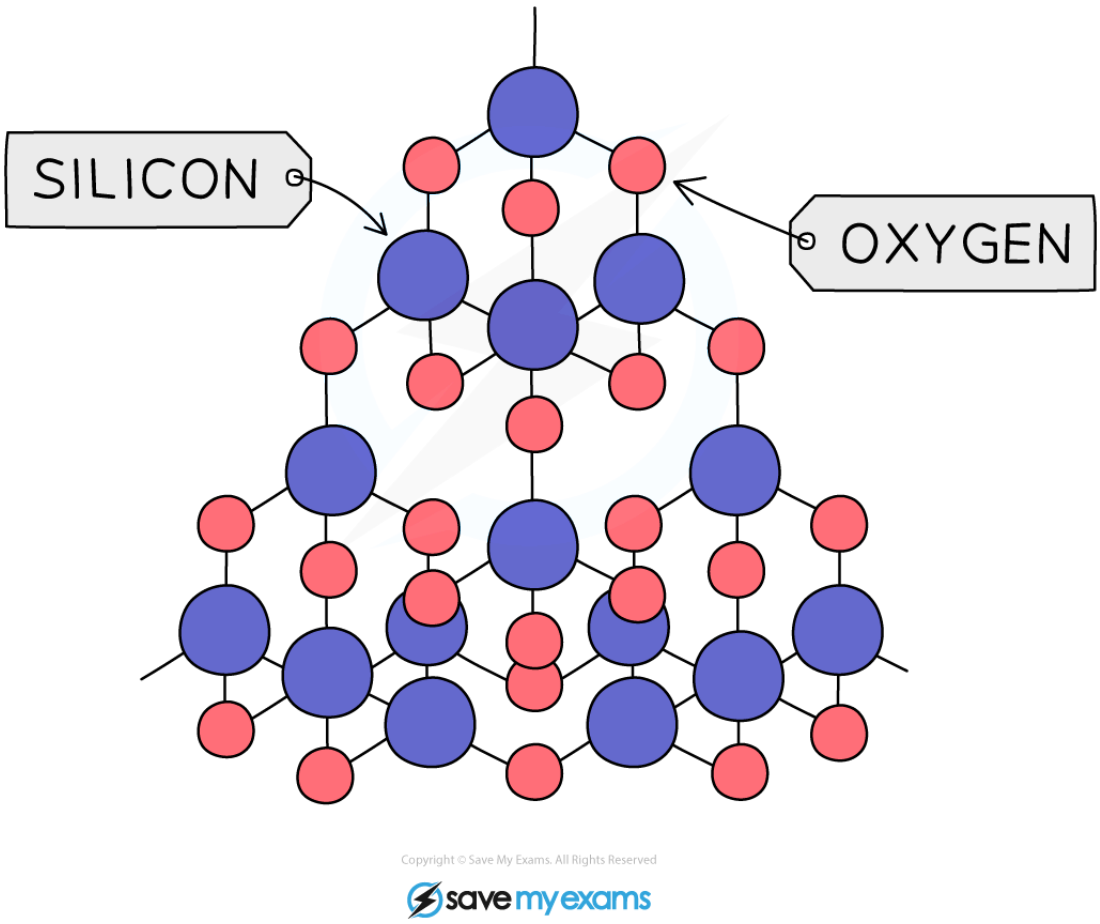
Diagram showing the structure and bonding arrangement in diamond

Silicon(IV) oxide

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Your notes

- Silicon(IV) oxide (also known as silicon dioxide or silica), SiO_2, is a macromolecular compound which occurs naturally as **sand** and **quartz**
- Each oxygen atom forms covalent bonds with **2** silicon atoms and each silicon atom in turn forms covalent bonds with **4** oxygen atoms
- A tetrahedron is formed with one silicon atom and four oxygen atoms, similar to diamond



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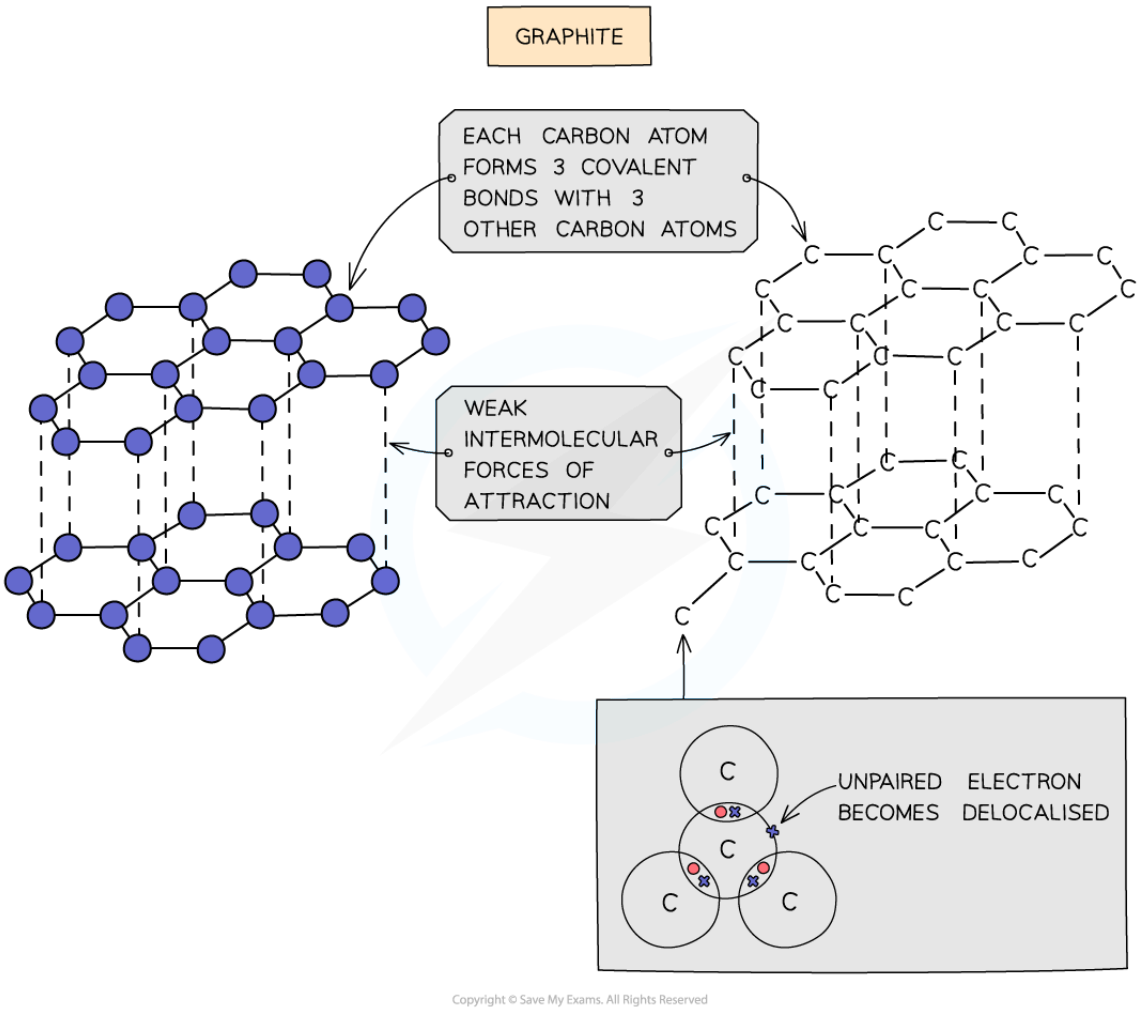
Diagram showing the structure of SiO_2 with the silicon atoms in blue and the oxygen atoms in red

Graphite

- Each carbon atom in graphite is bonded to **three** others forming **layers** of **hexagons**, leaving one free electron per carbon atom which becomes delocalised
- The covalent bonds within the layers are very strong, but the layers are attracted to each other by weak **intermolecular forces**

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Your notes



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The structure and bonding in graphite

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Uses of Graphite & Diamond


Your notes

Properties of Diamond

- Diamond has the following physical properties:
 - It does not conduct electricity
 - It has a very high melting point
 - It is extremely hard and dense

- All the outer shell electrons in carbon are held in the four covalent bonds around each carbon atom, so there are no freely moving charged particles to carry the current thus it cannot conduct electricity
- The four covalent bonds are very strong and extend in a giant lattice, so a very large amount of heat energy is needed to break the lattice thus it has a very high melting point
- Diamond is hard due to each carbon atom being covalently bonded to four other carbon atoms.
- It's hardness makes it very useful for purposes where extremely tough material is required
- Diamond is used in **jewellery** due to its sparkly appearance and as **cutting tools** as it is such a hard material
- The cutting edges of discs used to cut bricks and concrete are tipped with diamonds
- Heavy-duty drill bits and tooling equipment are also diamond-tipped

Exam Tip

Diamond is the hardest naturally occurring mineral, but it is by no means the strongest. Students often confuse **hard** with **strong**, thinking it is the opposites of weak. Diamonds are hard, but brittle – that is, they can be smashed fairly easily with a hammer. The opposite of saying a material is hard is to describe it as **soft**.

Properties of Graphite

- Each carbon atom is bonded to **three** others forming **layers** of hexagonal-shaped forms, leaving one free electron per carbon atom
- These free (delocalised) electrons exist **in between** the layers and are free to move through the structure and carry charge, hence graphite can conduct electricity
- The covalent bonds within the layers are very strong but the layers are connected to each other by **weak forces** only, hence the layers can **slide** over each other making graphite **slippery** and **smooth**
- Graphite thus:
 - Conducts electricity
 - Has a very high melting point
 - Is soft and slippery, less dense than diamond

- Graphite is used in **pencils** and as an industrial **lubricant**, in engines and in locks
- It is also used to make non-reactive **electrodes** for **electrolysis**

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Exam Tip

Don't confuse pencil lead with the metal lead – they have nothing in common. Pencil lead is actually graphite, and historical research suggests that in the past, lead miners sometimes confused the mineral galena (lead sulfide) with graphite; since the two looked similar they termed both minerals 'lead'. The word graphite derives from the Latin word 'grapho' meaning 'I write', so it is a well named mineral!



Your notes

Comparing Diamond & Silicon(IV) Oxide

- SiO_2 has lots of very strong covalent bonds and no intermolecular forces so it has similar properties to diamond
- It is very **hard**, has a very **high** boiling point, is insoluble in water and does not conduct electricity
- SiO_2 is cheap since it is available naturally and is used to make sandpaper and to line the inside of furnaces

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Cambridge O Level Chemistry

Your notes

2.7 Metallic Bonding

Contents

Metallic Bonding

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Metallic Bonding

Your notes

Metallic Bonding

- Metal atoms are held together strongly by metallic bonding in a **giant** metallic lattice
- Within the metallic lattice, the atoms lose the electrons from their outer shell and become positively charged ions
- The outer electrons no longer belong to a particular metal atom and are said to be **delocalised**
- They move freely between the positive metal ions like a 'sea of electrons'
- Metallic bonds are strong and are a result of the **electrostatic attraction** between the positive metal ions and the negatively charged delocalised electrons

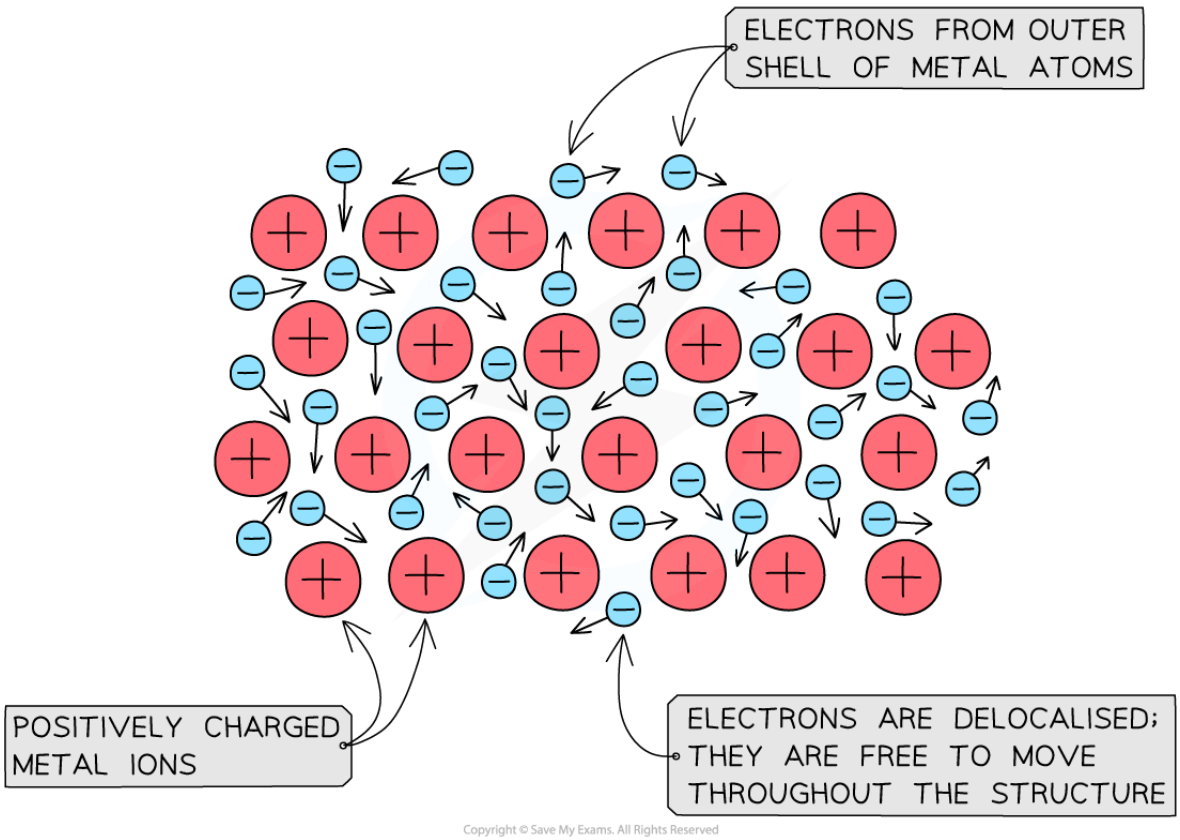


Diagram showing metallic lattice structure with delocalised electrons

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Properties of Metals

- Metals **conduct** electricity
 - There are **delocalised electrons** that are free to move through the structure and carry charge
 - Electrons entering one end of the metal cause a delocalised electron to displace itself from the other end
 - Hence electrons can flow so electricity is conducted
- Metals are **malleable** and **ductile**
 - Layers of positive ions can **slide** over one another and take up different positions
 - Metallic bonding is not disrupted as the outer electrons do not belong to any particular metal atom so the delocalised electrons will move with them
 - Metallic bonds are thus not broken and as a result metals are strong but **flexible**
 - They can be hammered and bent into different shapes or drawn into wires without breaking

Exam Tip

When explaining why metals can conduct electricity, be careful of the terminology you use. Don't get confused with ionic compounds. **Metals** can conduct electricity as they have free **electrons** that can carry charge whereas molten or aqueous **ionic** compounds can conduct electricity because they have free **ions** that can carry charge.

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Cambridge O Level Chemistry

Your notes

3.1 Formulae

Contents

- Formulae
- Empirical Formulae & Formulae of Ionic Compounds
- Writing Equations

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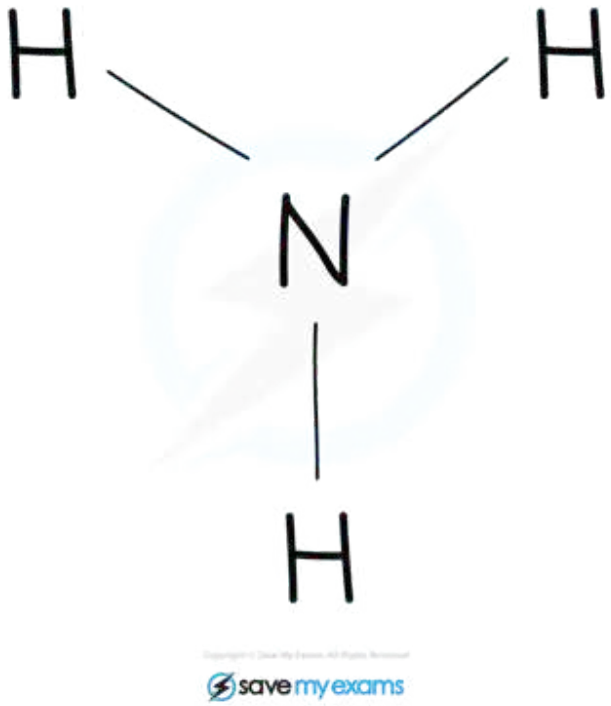
Formulae

Your notes

Molecular Formulae

Element symbols

- Each element is represented by its own unique symbol as seen on the Periodic Table
 - E.g. H is hydrogen
- Where a symbol contains two letters, the first one is always in **capital** letters and the other is **small**
 - E.g. sodium is Na, not NA
- Atoms combine together in **fixed ratios** that will give them full outer shells of electrons
- The chemical formula tells you the ratio of atoms
 - E.g. H_2O is a compound containing 2 hydrogen atoms which combine with 1 oxygen atom
- The chemical formula can be deduced from the relative number of atoms present
 - E.g. If a molecule contains 3 atoms of hydrogen and 1 atom of nitrogen then the formula would be NH_3
- Diagrams or models can also be used to represent the chemical formula



The ammonia molecule consists of a central nitrogen atom bonded to three hydrogen atoms

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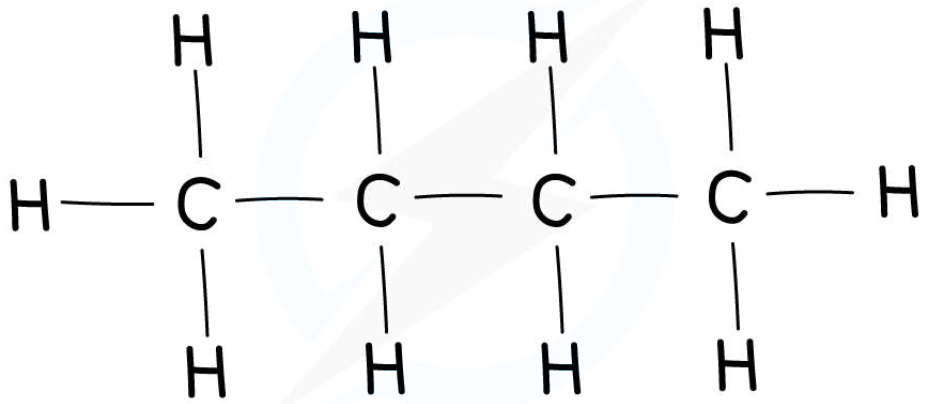
Your notes

Chemical formulae

- The **structural** formula tells you the way in which the atoms in a particular molecule are bonded
 - This can be done by either a **diagram** (displayed formula) or **written** (simplified structural formula)
- The **molecular** formula tells you the actual number of atoms of each element in one molecule of the compound or element
 - E.g. H_2 has 2 hydrogen atoms, HCl has 1 hydrogen atom and 1 chlorine atom

Example: Butane

- **Structural formula (displayed)**



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- **Structural formula (simplified)**

$CH_3CH_2CH_2CH_3$

- **Molecular formula**

C_4H_{10}

- **Empirical formula**

C_2H_5

Deducing formulae by Combining power(valency)

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- The concept of valency is used to deduce the formulae of compounds (either molecular compounds or ionic compounds)
- Valency or combining power tells you how many bonds an atom can make with another atom or how many electrons its atoms lose, gain or share, to form a compound
 - E.g. carbon is in Group IV so a single carbon atom can make 4 single bonds or 2 double bonds
- The following valencies apply to elements in each group:

GROUP	VALENCY
I	1
II	2
III	3
IV	4
V	3
VI	2
VII	1
VIII	0


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Worked example

What is the formula of aluminium sulfide?

Answer:

- We can use the combining power (valency) of each atom to work out a formula

Al **S** **3** **2**	Write out the symbols of each element and write their combining powers underneath
Al **S**  **3** **2**	The formula is then calculated by cross multiplying each atom with the number opposite
Al_2S_3	The formula for aluminium sulfide is Al_2S_3

Your notes

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Empirical Formulae & Formulae of Ionic Compounds


Your notes

Empirical Formulae

The **molecular formula** is the formula that shows the **number** and **type** of each atom in a molecule

- E.g. the molecular formula of ethanoic acid is $C_2H_4O_2$
- The **empirical formula** is the simplest whole number ratio of the atoms of each element present in one molecule or formula unit of the compound
 - E.g. the empirical formula of ethanoic acid is CH_2O
- **Organic molecules**, such as ethanoic acid, often have **different** empirical and molecular formulae
- The formula of an **ionic compound** is always an **empirical formula**

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Deducing Formulae of Ionic Compounds


Your notes

- The formulae of these compounds can be calculated if you know the charge on the ions
- The Periodic Table can help work out the charge on many elements:
 - Group I elements form ions with a 1+ charge
 - Group II elements form ions with a 2+ charge
 - Group III elements form ions with a 3+ charge
 - Group V elements form ions with a 3– charge
 - Group VI elements form ions with a 2– charge
 - Group VII elements form ions with a 1– charge
- Below are some other common ions and their charges
- Note that a Roman numeral next to the element tells you the charge on the ion, e.g. copper(II) ions have a charge 2+
- There are several common compound ions included in the table
 - Some chemists call these polyatomic ions

Common Ions & Their Charges Table

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Your notes

Ion	Formula and charge
Iron(II)	Fe^{2+}
Copper(II)	Cu^{2+}
Chromium(III)	Cr^{3+}
Ammonium	NH_4^+
Hydroxide	OH^-
Nitrate	NO_3^-
Sulfate	SO_4^{2-}
Carbonate	CO_3^{2-}
Hydrogen carbonate	HCO_3^-

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- The overall sum of the charges of an ionic compound should be 0
- You therefore need to work out the ratio of the ions to ensure this is the case
- When you write the formula of a compound ion it is necessary to use brackets around the compound ion where more than one of that ion is needed in the formula
 - For example copper(II) hydroxide is $Cu(OH)_2$

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Worked example

What is the formula of?

1. sodium bromide
2. aluminium fluoride
3. aluminium oxide
4. magnesium nitrate
5. ammonium sulfate

Answer 1

Symbol	Na Br
Ion charge	1+ 1–
Balance the number of ions	1 sodium ion is needed for each bromide ion
Ratio of ions	1:1
Formula	NaBr

Answer 2

Symbol	Al F
Ion charge	3+ 1–
Balance the number of ions	3 fluoride ions are needed for each aluminium ion
Ratio of ions	1:3
Formula	AlF_3

Answer 3

Symbol	Al O
Ion charge	3+ 2–
Balance the number of ions	2 aluminium ions are needed for 3 oxide ions
Ratio of ions	2:3
Formula	Al_2O_3

Answer 4

Symbol	Mg^{2+} NO_3^-
Ion charge	2+ 1–

Your notes

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Balance the number of ions	2 nitrate ions are needed for each magnesium ion
Ratio of ions	1:2
Formula	$Mg(NO_3)_2$

Answer 5

Symbol	NH_4^+ SO_4^{2-}
Ion charge	1+ 2–
Balance the number of ions	2 ammonium ions are needed for each sulfate ion
Ratio of ions	2:1
Formula	$(NH_4)_2SO_4$

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Writing Equations

Your notes

Writing Word Equations, Symbol Equations & Ionic Equations

Word equations

- These show the reactants and products of a chemical reaction using their full chemical names
- The arrow (which is spoken as "goes to" or "produces") implies the conversion of reactants into products
- Reaction conditions or the name of a catalyst can be written above the arrow
- An example of a word equation for neutralisation is:

sodium hydroxide + hydrochloric acid → sodium chloride + water

- The reactants are sodium hydroxide and hydrochloric acid
- The products are sodium chloride and water

Names of compounds

- For compounds consisting of 2 atoms:
 - If one is a **metal** and the other a **non-metal**, then the name of the metal atom comes first and the ending of the second atom is replaced by adding -**ide**
 - E.g. NaCl which contains sodium and chlorine thus becomes sodium chlor**ide**
 - If both atoms are **non-metals** and one of those is **hydrogen**, then hydrogen comes first
 - E.g. Hydrogen and chlorine combined is called hydrogen chlor**ide**
- For other combinations of non-metals as a general rule, the element that has a lower group number comes first in the name
 - E.g. carbon and oxygen combine to form CO_2 which is carbon dioxide since carbon is in Group 4 and oxygen in Group 6
- For compounds that contain certain groups of atoms:
 - There are **common groups** of atoms which occur regularly in chemistry
 - Examples include the carbonate ion (CO_3^{2-}), sulfate ion (SO_4^{2-}), hydroxide ion (OH^-) and the nitrate ion (NO_3^-)
 - When these ions form a compound with a metal atom, the name of the **metal** comes **first**
 - E.g. KOH is potassium hydroxide, $CaCO_3$ is calcium carbonate

Writing and balancing chemical equations

- Chemical equations use the chemical symbols of each reactant and product
- When balancing equations, there needs to be the **same number of atoms** of each element on either side of the equation
- The following non-metals must be written as diatomic molecules (i.e. molecules that contain two atoms): H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2

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- Work across the equation from left to right, checking one element after another
- If there is a group of atoms, for example a nitrate group (NO_3^-) that has not changed from one side to the other, then count the whole group as one entity rather than counting the individual atoms.
 - Examples of chemical equations:
 - Acid-base neutralisation reaction:
 $NaOH\,(aq) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$
 - Redox reaction:
 $2Fe_2O_3\,(s) + 3C\,(s) \longrightarrow 4Fe\,(s) + 3CO_2\,(g)$
 - In each equation there are equal numbers of each atom on either side of the reaction arrow so the equations are balanced
- The best approach is to practice lot of examples of balancing equations
- By trial and error change the coefficients (multipliers) in front of the formulae, one by one checking the result on the other side
- Balance elements that appear on their own, last in the process


Your notes

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Worked example

Example 1

Balance the following equation:

aluminium + copper(II)oxide ⟶ aluminium oxide + copper

Unbalanced symbol equation:

$$Al + CuO \longrightarrow Al_2O_3 + Cu$$

$$\mathbf{2}Al + CuO \longrightarrow Al_2O_3 + Cu$$

STEP 1: ONLY 1 Al ATOM ON THE LEFT AND 2 ON THE RIGHT SO IF YOU END UP WITH 2, YOU MUST START WITH 2. CHANGE Al TO 2 Al

$$2Al + \mathbf{3}CuO \longrightarrow Al_2O_3 + Cu$$

STEP 2: ONLY 1 O ATOM ON THE LEFT AND 3 ON THE RIGHT SO IF YOU END UP WITH 3, YOU MUST START WITH 3. TO ACHIEVE THIS, IT MUST BE 3CuO

$$2Al + 3CuO \longrightarrow Al_2O_3 + \mathbf{3}Cu$$

STEP 3: THERE ARE 3 Cu ATOMS ON THE LEFT AND 1 ON THE RIGHT. THE ONLY WAY OF ACHIEVING 3 ON THE RIGHT IS TO HAVE 3 Cu

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Your notes

Worked example

Example 2

Balance the following equation:

magnesium oxide + nitric acid ⟶ magnesium nitrate + water

Unbalanced symbol equation:

$$MgO + HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$$

$MgO + HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

STEP 1: THERE IS 1 Mg ATOM ON THE LEFT AND 1 ON THE RIGHT SO THERE ARE EQUAL NUMBERS ON BOTH SIDES SO THESE ARE KEPT THE SAME

$MgO + \mathbf{2}HNO_3 \longrightarrow Mg(NO_3)_2 + H_2O$

STEP 2: THERE IS 1 O ATOM ON THE LEFT AND 1 ON THE RIGHT SO THERE IS AN EQUAL NUMBER ON BOTH SIDES. REMEMBER THAT YOU ARE COUNTING THE NITRATE GROUP AS A SEPARATE GROUP, SO DO NOT COUNT THE OXYGEN ATOMS IN THIS GROUP. BECAUSE THERE IS 1 NITRATE GROUP ON THE LEFT, BUT 2 NITRATE GROUPS ON THE RIGHT, THESE NEED TO BE BALANCED. SO, ADD A 2 IN FRONT OF THE HNO_3

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State symbols

- State symbols are written after each **formula** in chemical equations to show which physical state each substance is in
- Brackets are used and they are not usually subscripted although you may come across them written in this way
- Aqueous should remind you of the word 'aqua' and means the substance is **dissolved** in **water**
 - In other words it is a solution

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SOLID	LIQUID	GAS	AQUEOUS
(s)	(l)	(g)	(aq)

The four state symbols show the physical state of substances at normal conditions

- Symbol equations should be included when writing chemical equations.
- An example of a reaction with state symbols is the reaction of copper carbonate with hydrochloric acid:

$$CuCO_3\,(s) + 2HCl\,(aq) \rightarrow CuCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

- Sometimes it can be hard to know what the correct state symbol is and we have to look for clues in the identity of substances in a reaction
- Generally, unless they are in a solution:
 - Metal compounds will always be solid, although there are a few exceptions
 - Ionic compounds will usually be solids
- Non-metal compounds could be solids, liquids or gases, so it depends on chemical structure
- Precipitates formed in solution count as solids
- In the worked examples above the final equations with the state symbols would be
 - $2Al\,(s) + 3CuO\,(s) \rightarrow Al_2O_3\,(s) + 3Cu\,(s)$
 - $MgO\,(s) + 2HNO_3\,(aq) \rightarrow Mg(NO_3)_2\,(aq) + H_2O\,(l)$

Exam Tip

Be careful when writing the state symbol of solutions of liquids. For example, ethanol, or common alcohol, is a liquid at room temperature, so if it is pure alcohol then you would be using (l) as the state symbol; most of the time alcohol is used as a solution in water so (aq) is symbol to use.

Balancing Ionic Equations

- In aqueous solutions ionic compounds **dissociate** into their ions, meaning they separate into the component ions that formed them
- E.g. hydrochloric acid and potassium hydroxide dissociate as follows:

$$HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-(aq)$$

$$KOH\,(aq) \rightarrow K^+\,(aq) + OH^-\,(aq)$$

- It is important that you can recognise common ionic compounds and their constituent ions
- These include:
 - Acids such as HCl and H_2SO_4

Your notes

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- Group I and Group II hydroxides e.g. sodium hydroxide
- Soluble salts e.g. potassium sulfate, sodium chloride

- Follow the example below to write ionic equations

Your notes

Worked example

Write the ionic equation for the reaction of aqueous chlorine and aqueous potassium iodide.4

Answer:

- **Step 1:** Write out the full balanced equation:

$$2KI\,(aq) + Cl_2\,(aq) \rightarrow 2KCl\,(aq) + I_2\,(aq)$$

- **Step 2:** Identify the ionic substances and write down the ions separately

$$2K^+\,(aq) + 2I^-\,(aq) + Cl_2\,(aq) \rightarrow 2K^+\,(aq) + 2Cl^-\,(aq) + I_2\,(aq)$$

- **Step 3:** Rewrite the equation eliminating the ions which appear on both sides of the equation (spectator ions) which in this case are the K^+ ions:

$$2I^-\,(aq) + Cl_2\,(aq) \rightarrow 2Cl^-\,(aq) + I_2\,(aq)$$

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Deducing Symbol Equations

- For some reactions, you will not be given the unbalanced equation but you will be expected to use your knowledge learnt throughout the course to know or deduce the formula of compounds and then balance the equations

Your notes

Worked example

Aluminium burns in chlorine to form the white solid, aluminium chloride. Write the balanced symbol equation, including state symbols, for the reaction.

Answer:

Step 1: Work out the formula and state symbols of the reactants and products to construct an unbalanced symbol equation:

- Aluminium is a solid metal, like other pure metals, it is an element so its formula is the same as its chemical symbol: **Al (s)**
- From your knowledge of Group VII elements, you should know that chlorine is a gas that exists as a diatomic molecule: **Cl_2 (g)**
- Aluminum chloride is a solid - this information is given in the question as you would not be expected to know this. Its formula is deduced from the charges on the ions present:
- Aluminium has a 3+ charge and chloride ions have a 1– charge, therefore for the compound to be neutral, 3 chloride ions are needed for every 1 aluminium ion: **$AlCl_3$ (s)**
- The unbalanced symbol equation is thus:

$$Al\,(s) + Cl_2\,(g) \rightarrow AlCl_3\,(s)$$

Step 2: Balance the equation:

- Make the number of Cl on the RHS an even number by adding a 2 in front of $AlCl_3$:

$$Al\,(s) + Cl_2\,(g) \rightarrow \mathbf{2}AlCl_3\,(s)$$

- This gives 6 Cl on the RHS so now balance the number of Cl on the LHS by adding a 3 in front of Cl_2:

$$Al\,(s) + \mathbf{3}Cl_2\,(g) \rightarrow \mathbf{2}AlCl_3\,(s)$$

- Finally, there are now 2 Al on the RHS but only 1 on the LHS, so add a 2 in front of the Al on the LHS:

$$\mathbf{2Al\,(s) + 3Cl_2\,(g) \rightarrow 2AlCl_3\,(s)}$$

Exam Tip

When balancing equations you **cannot** change any of the formulae, only the amount of each atom or molecule. This is done by changing the numbers that go in **front** of each chemical species.

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Cambridge O Level Chemistry

Your notes

3.2 Relative Masses of Atoms & Molecules

Contents

* Ar & Mr

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Ar & Mr

Your notes

Relative Masses

Relative Atomic Mass

- The symbol for the relative atomic mass is A_r
- The relative atomic mass for each element can be found in the Periodic Table along with the atomic number
- The relative atomic mass is shown underneath the atomic symbol and is larger than the atomic number (except for hydrogen where they are the same)
- Atoms are too small to accurately weigh but scientists needed a way to compare the masses of atoms
- The carbon-12 is used as the standard atom and has a fixed mass of 12 units
- It is against this atom which the masses of all other atoms are compared
- Relative atomic mass (A_r) can therefore be defined as:
 - **The average mass of the isotopes of an element compared to 1/12th of the mass of an atom of ^{12}C**
- The relative atomic mass of carbon is 12
 - The relative atomic mass of magnesium is 24 which means that magnesium is twice as heavy as carbon
 - The relative atomic mass of hydrogen is 1 which means it has one twelfth the mass of one carbon-12 atom

Relative molecular (formula) mass

- The symbol for the relative molecular mass is M_r and it refers to the **total mass** of the molecule
- To calculate the M_r of a substance, you have to add up the relative atomic masses of all the atoms present in the formula
- **Relative formula mass** is used when referring to the total mass of an **ionic** compound

Relative Formula Mass Calculations Table

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Your notes

SUBSTANCE	ATOMS PRESENT	M_r
HYDROGEN (H_2)	$2 \times H$	$(2 \times 1) = 2$
WATER (H_2O)	$(2 \times H) + (1 \times O)$	$(2 \times 1) + 16 = 18$
POTASSIUM CARBONATE (K_2CO_3)	$(2 \times K) + (1 \times C) + (3 \times O)$	$(2 \times 39) + 12 + (3 \times 16) = 138$
CALCIUM HYDROXIDE ($Ca(OH)_2$)	$(1 \times Ca) + (2 \times O) + (2 \times H)$	$40 + (2 \times 16) + (2 \times 1) = 74$
AMMONIUM SULFATE ($(NH_4)_2SO_4$)	$(2 \times N) + (8 \times H) + (1 \times S) + (4 \times O)$	$(2 \times 14) + (8 \times 1) + 32 + (4 \times 16) = 132$

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Reacting masses

- The Law of Conservation of mass tells us that mass cannot be created or destroyed
- In a chemical reaction, the total mass of reactants equals the total mass of the products
- We can use this, along with relative atomic/formula masses to perform calculations to identify the quantities of reactants or products involved in a chemical reaction
- **Example**:

$$2Ca + O_2 \rightarrow 2CaO$$

- Relative atomic masses: Ca = 40; O = 16
- Using the balanced symbol equation shows that 2 x 40 = 80 units of mass of calcium react with 2 x 16 = 32 units of mass of oxygen to form 2 x (40 + 16) = 112 units of mass of CaO:

$$2Ca + O_2 \rightarrow 2CaO$$

$$80 + 32 = 112$$

- The ratio of the mass of calcium and oxygen reacting will always be the same, regardless of the units
 - E.g. 80 g of calcium will react with 32 g of oxygen to form 112 g of calcium oxide
 - Or, 40 tonnes of calcium will react in excess oxygen to form 56 tonnes of calcium oxide

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Worked example

Calculate the mass of carbon dioxide produced when 32 g of methane, CH_4, reacts completely in excess oxygen:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Relative atomic masses, A_r: H = 1; C = 12; O = 16

Answer

- In terms of relative mass the equation is:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$12 + (4 \times 1) + 2 \times (2 \times 16) \rightarrow 12 + (2 \times 16) + 2 \times (2 \times 1 + 16)$

$16 + 64 \rightarrow 44 + 36$

- So 16 g of methane would react in excess oxygen to form 44 g of carbon dioxide
- Therefore, 32 g of methane would produce 44 x 2 = **88 g** of carbon dioxide

Your notes

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Cambridge O Level Chemistry

Your notes

3.3 The Mole & the Avogadro Constant

Contents

- The Mole
- Linking Moles, Mass & Mr
- Molar Gas Volume
- Calculating Concentration
- Reacting Masses
- Titration Calculations
- Empirical & Molecular Formula
- Percentage Yield & Purity

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The Mole


Your notes

The Mole & the Avogadro Constant

The Mole & Avogadro's Constant

- Chemical amounts are measured in moles
- The mole, symbol mol, is the SI unit of **amount of substance**
- One mole of a substance contains the same number of the stated particles, atoms, molecules, or ions as one mole of any other substance
- One mole contains **6.02×10^{23}** particles (e.g. atoms, ions, molecules); this number is known as the **Avogadro constant**
- For example:
 - One mole of sodium (Na) contains 6.02×10^{23} atoms of sodium
 - One mole of hydrogen (H_2) contains 6.02×10^{23} molecules of hydrogen
 - One mole of sodium chloride (NaCl) contains 6.02×10^{23} formula units of sodium chloride
- The mass of 1 mole of a substance is known as the **molar mass**
- For an element, it is the same as the **relative atomic mass** written in grams
- For a compound, it is the same as the **relative formula mass** or **relative molecular mass** in grams

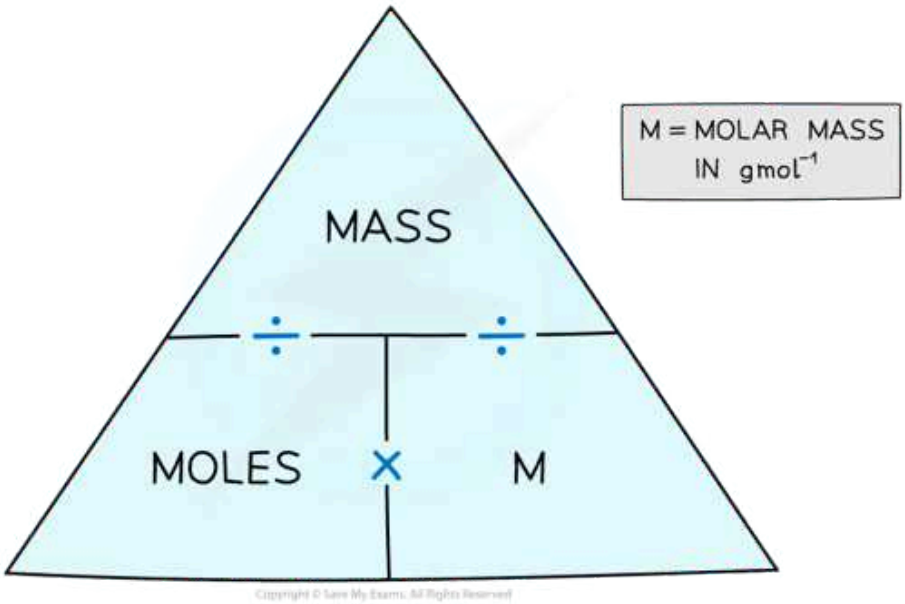
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Linking Moles, Mass & Mr

Your notes

Linking Moles, Mass & Mr

- Although elements and chemicals react with each other in **molar ratios**, in the laboratory we use digital balances and grams to measure quantities of chemicals as it is impractical to try and measure out moles
- Therefore we have to be able to convert between moles and grams
- We can use the following formula to convert between moles, mass in grams and the molar mass
- The mass of 1 mole of a substance is known as the **molar mass**
- For an element, it is the same as the **relative atomic mass** written in grams
- For a compound it is the same as the **relative formula mass** or **relative molecular mass** in grams



Formula triangle for moles, mass and molar mass

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Your notes

Worked example

What is the mass of 0.250 moles of zinc?

Answer

- From the Periodic Table, the relative atomic mass of Zn is 65.4
- So, the molar mass is 65.4 g/mol
- The mass is calculated by **moles x molar mass**
- This comes to 0.250 mol x 65.4 g/mol = **16.35 g**

Worked example

How many moles are in 2.64 g of sucrose, $C_{12}H_{22}O_{11}$ (M_r = 342.3)?

Answer

- The molar mass of sucrose is 342.3 g/mol
- The number of moles is found by **mass ÷ molar mass**
- This comes to 2.64 g ÷ 342.3 g/mol = **7.71×10^{-3} mol**

Worked example

In 15.7 g of water (M_r = 18):

a. How many molecules are there?
b. How many atoms are there?

Answer – part a

- The molar mass of water is 18 g/mol
- The number of moles is found by **mass ÷ molar mass**
- This comes to 15.7 g ÷ 18 g/mol = **0.872 mol**
- There are 6.02×10^{23} molecules of water in 1 mole of water
- In 0.872 moles of water, there are $6.02 \times 10^{23} \times 0.872$ = **5.25×10^{23}** molecules

Answer – part b

- In each molecule of water there are 3 atoms (2 hydrogen atoms, one oxygen atom)
- The number of atoms in 15.7 g = $3 \times 5.25 \times 10^{23}$ = **1.58×10^{24}** atoms

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Exam Tip

Always show your workings in calculations as its easier to check for errors and you may pick up credit if you get the final answer wrong.



Your notes

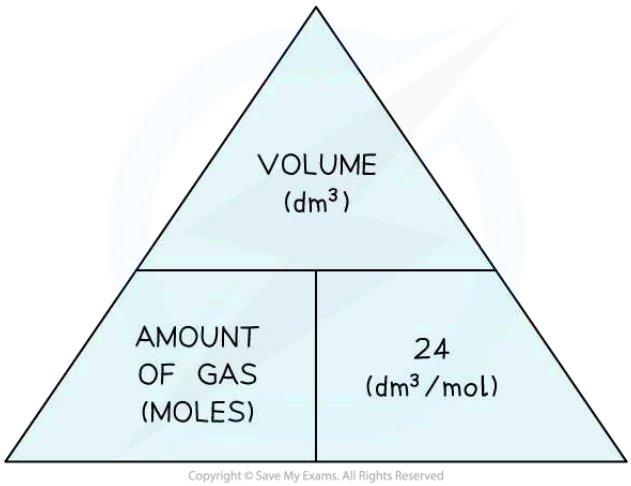
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Molar Gas Volume

Your notes

Molar Gas Volume

- **Avogadro's Law** states that at the same conditions of **temperature** and **pressure**, equal amounts of gases occupy the **same volume** of space
- At room temperature and pressure, the volume occupied by one mole of any gas was found to be **24 dm^3** or **24,000 cm^3**
- This is known as the **molar gas volume at RTP**
- RTP stands for "room temperature and pressure" and the conditions are **20 °C** and **1 atmosphere** (atm)
- From the molar gas volume the following formula triangle can be derived:



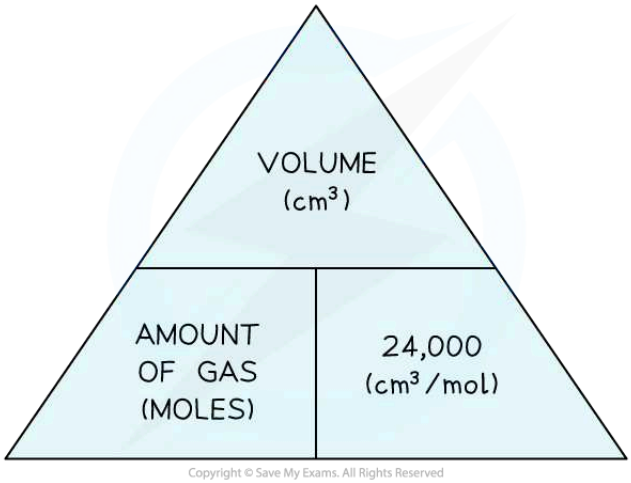
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Formula triangle showing the relationship between moles of gas, volume in dm^3 and the molar volume

- If the volume is given in cm^3 instead of dm^3, then divide by 24,000 instead of 24:

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Your notes



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Formula triangle showing the relationship between moles of gas, volume in cm^3 and the molar volume

- The formula can be used to calculate the number of moles of gases from a given volume or vice versa
- Simply cover the one you want and the triangle tells you what to do

To find the volume of a gas

Volume = Moles x Molar Volume

Examples of Converting Moles into Volumes Table

NAME OF GAS	AMOUNT OF GAS	VOLUME OF GAS
HYDROGEN	3 mol	(3 x 24) = 72 dm^3
CARBON DIOXIDE	0.25 mol	(0.25 x 24) = 6 dm^3
OXYGEN	5.4 mol	(5.4 x 24,000) = 129,600 cm^3
AMMONIA	0.02 mol	(0.02 x 24) = 0.48 dm^3

To find the moles of a gas

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Moles = Volume ÷ Molar Volume

Examples of Converting Volumes into Moles Table

NAME OF GAS	VOLUME OF GAS	MOLES OF GAS
METHANE	225.6 dm^3	(225.6 : 24) = 9.4 mol
CARBON MONOXIDE	7.2 dm^3	(7.2 : 24) = 0.3 mol
SULFUR DIOXIDE	960 dm^3	(960 : 24) = 40 mol
OXYGEN	1200 cm^3	(1200 : 24000) = 0.05 mol

Your notes

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Calculating Gas Volumes

Your notes

- You may be asked to calculate the volume of a gas from a given amount stated in grams instead of moles
- To answer these type of questions you must first convert grams to moles and then calculate the volume.

Worked example

What is the volume of 154 g of nitrogen gas at RTP?

Answer

Step	Working out
1. Calculate the moles of nitrogen, remembering that nitrogen gas is diatomic	Mr of N_2 = 2 × 14 = 28g Moles of N_2 = $\frac{154}{28}$ = 5.5 moles
2. Calculate the volume of gas from the molar volume equation	5.5 × 24 = 132 dm^3

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- A second style of gas calculation involves calculating the volumes of gaseous reactants and products from a balanced equation and a given volume of a gaseous reactant or product
- These problems are straightforward as you are applying **Avogadro's Law**, so the moles (and coefficients) in equations are in the same ratio as the gas volumes

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Worked example

The complete combustion of propane gives carbon dioxide and water vapour as the products

$$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)$$

Determine the volume of oxygen needed to react with 150 cm^3 of propane and the total volume of the gaseous products.

Answer

- The balanced equation shows that 5 moles of oxygen are needed to completely react with 1 mole of propane
- Therefore the volume of oxygen needed would be = 5 moles x 150 cm^3 = **750 cm^3**
- The total number of moles of gaseous products is = 3 + 4 = 7 moles
- The total volume of gaseous products would be = 7 moles x 150 cm^3 = **1050 cm^3**

Exam Tip

Make sure you use the correct units as asked by the question when working through reacting gas volume questions.

Your notes

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Calculating Concentration


Your notes

Units of Concentration

- A solid substance that dissolves in a liquid is called a solute, the liquid is called a solvent and the two when mixed together form a solution
- Most chemical reactions occur between solutes which are dissolved in solvents, such as water or an organic solvent
- Concentration simply refers to the amount of solute there is in a specific volume of the solvent
- The greater the amount of solute in a given volume, the greater the concentration
- The amount of solute can be expressed in grams or moles
- Typically, concentration is expressed in terms of the amount of substance per dm^3, therefore the units of concentration are either: **g/dm^3** or **mol/dm^3**

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Calculating Concentration

Your notes

- **Concentration** refers to the amount of solute there is in a specific volume of the solvent
- A general formula to calculate the concentration in g/dm^3 is:

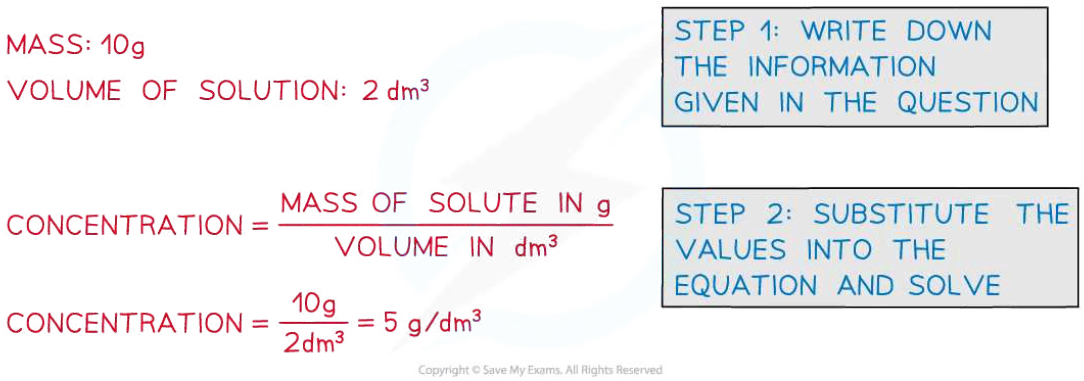
$$\text{concentration } (\text{g/dm}^3) = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3)}$$

- Concentration can be measured in grams per cubic decimetre
- 1 decimetre cubed (dm^3) = 1000 cm^3
 - 1 decimetre cubed (dm^3) is the same as 1 litre
- You may be given data in a question which needs to be converted from cm^3 to dm^3 or the other way around
 - To go from cm^3 to dm^3:
 - Divide by 1000
 - To go from dm^3 to cm^3:
 - Multiply by 1000

Worked example

A student dissolved 10 g of sodium hydroxide, NaOH, in 2 dm^3 of distilled water. Calculate the concentration of the solution in g/dm^3.

Answer



Calculating Concentration using Moles

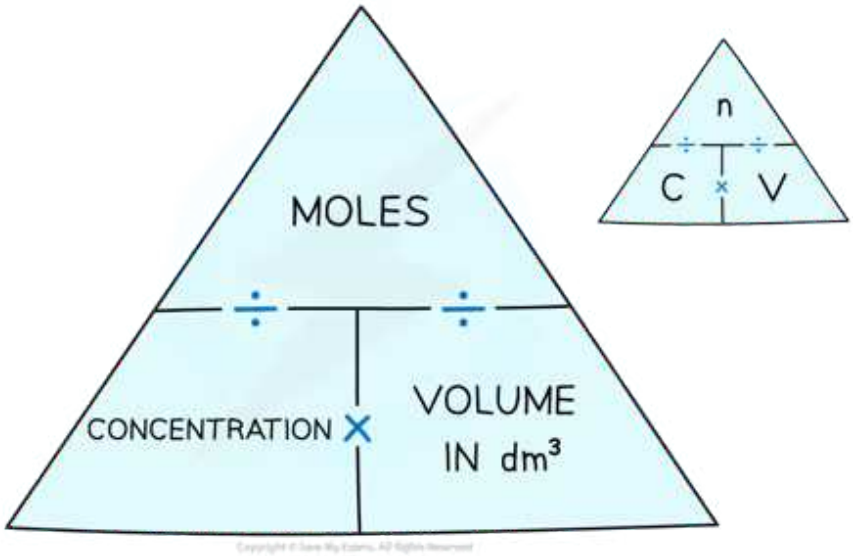
- It is more useful to a chemist to express concentration in terms of moles per unit volume rather than mass per unit volume
- Concentration can therefore be expressed in moles per decimetre cubed and calculated using the following equation:

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$$\text{concentration } (\text{mol/dm}^3) = \frac{\text{number of moles of solute (mol)}}{\text{volume of solution (dm}^3)}$$

Your notes

- We can modify the concentration formula to include moles
 - The units in the answer can be written as mol/dm^3 (this can also be written as mol dm^{-3})
- You may have to convert from g/dm^3 into mol/dm^3 and vice versa depending on the question
 - To go from g/dm^3 to mol/dm^3
 - Divide by the molar mass in grams
 - To go from mol/dm^3 to g/dm^3
 - Multiply by the molar mass in grams
- Some students find formula triangles help them to understand the relationship:



The concentration–moles formula triangle can help you solve these problems

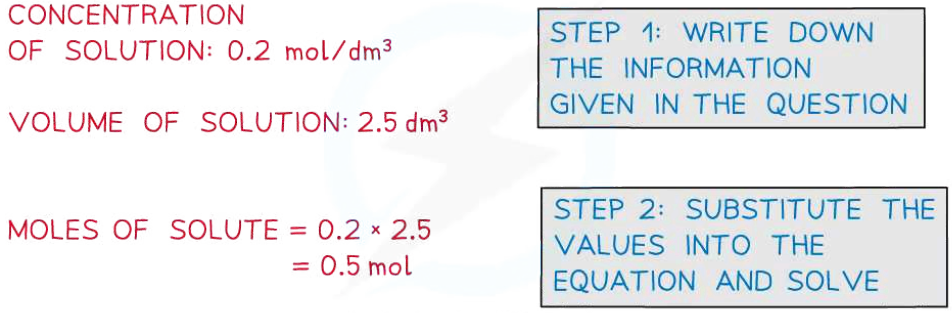
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Your notes

Worked example

Calculate the amount of solute, in moles, present in 2.5 dm^3 of a solution whose concentration is 0.2 mol/dm^3.

Answer



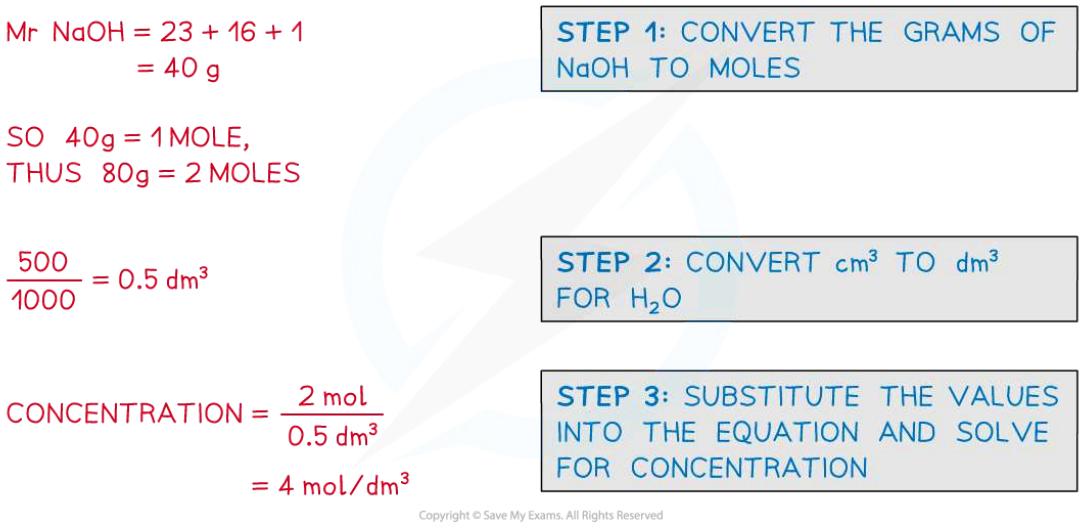
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Worked example

Calculate the concentration of a solution of sodium hydroxide, NaOH, in mol/dm^3, when 80 g is dissolved in 500 cm^3 of water.

Relative atomic masses, A_r: Na= 23; H= 1; O= 16

Answer



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Your notes

Worked example

25.0 cm^3 of 0.050 mol/dm^3 sodium carbonate was completely neutralised by 20.00 cm^3 of dilute hydrochloric acid. Calculate the concentration in mol/dm^3 of the hydrochloric acid.

Answer

AMOUNT OF Na_2CO_3 = (25.0 × 0.050) ÷ 1000
= 0.00125 mol

STEP 1: CALCULATE THE AMOUNT, IN MOLES, OF SODIUM CARBONATE REACTED BY REARRANGING THE EQUATION FOR AMOUNT OF SUBSTANCE (mol) AND DIVIDING BY 1000 TO CONVERT cm^3 TO dm^3

1 mol OF Na_2CO_3 REACTS WITH 2 mol OF HCl, SO THE MOLAR RATIO IS 1 : 2

THEREFORE 0.00125 moles OF Na_2CO_3 REACT WITH 0.00250 moles OF HCl

STEP 2: CALCULATE THE AMOUNT, IN MOLES, OF HYDROCHLORIC ACID REACTED

1 dm^3 = 1000cm^3
VOLUME OF HCl = 20 ÷ 1000
= 0.0200 dm^3

CONCENTRATION HCl (mol/dm^3) =
0.00250 ÷ 0.0200 = 0.125 mol/dm^3

STEP 3: CALCULATE THE CONCENTRATION, IN mol/dm^3 OF THE HYDROCHLORIC ACID

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Exam Tip

Remember to always convert the units from cm^3 to dm^3 by dividing by 1000.

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Reacting Masses

Your notes

Reacting Masses

- Chemical equations can be used to calculate the **moles** or **masses** of reactants and products
- To do this, information given in the question is used to find the amount in moles of the substances being considered
- Then, the **ratio** between the substances is identified using the balanced chemical equation
- Once the moles have been determined they can then be converted into grams using the relative atomic or relative formula masses

Worked example

Calculate the mass of magnesium oxide that can be made by completely burning 6.0 g of magnesium in oxygen in the following reaction:

$$2Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$$

Relative formula masses (M_r): Mg = 24; MgO = 40

$$\text{MOLES} = \frac{\text{MASS}}{\text{MOLAR MASS}}$$

$$\text{MOLES} = \frac{6}{24} = 0.25$$

STEP 1: CALCULATE THE MOLES OF MAGNESIUM USED IN THE REACTION

2Mg : 2MgO

RATIO IS THUS 1 : 1

SO 0.25 MOLES Mg PRODUCES 0.25 MOLES MgO

STEP 2: FIND THE RATIO OF Mg TO MgO USING THE MOLAR RATIO FROM THE BALANCED EQUATION

MASS = MOLES × MOLAR MASS

MASS = 0.25 × 40 = 10g

STEP 3: CALCULATE THE MASS OF MgO

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Your notes

Worked example

Calculate the mass of aluminium, in tonnes, that can be produced from 51 tonnes of aluminium oxide. The equation for the reaction is:

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

Relative formula masses (M_r): Al = 27; Al_2O_3 = 102

MASS OF AL_2O_3 IN g = 51×10^6
= 51,000,000 g

STEP 1: CALCULATE THE MOLES OF ALUMINIUM OXIDE USED

$$\text{MOLES} = \frac{\text{MASS}}{\text{MOLAR MASS}}$$

$$\text{MOLES} = \frac{51{,}000{,}000}{102} = 500{,}000$$

$2Al_2O_3 : 4Al$

RATIO IS THUS 1 : 2

SO 500,000 MOLES Al_2O_3 PRODUCES 1,000,000 MOLES Al

STEP 2: FIND THE RATIO OF Al_2O_3 TO Al USING THE MOLAR RATIO FROM BALANCED EQUATION

MASS = MOLES × MOLAR MASS

MASS = 1,000,000 × 27
= 27,000,000 g

STEP 3: CALCULATE THE MASS OF Al

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Exam Tip

Remember molar ratio of a balanced equation gives you the ratio of the amounts of each substance in the reaction.

Limiting Reactants

- A chemical reaction stops when one of the reactants is used up

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- The reactant that is used up first is the **limiting reactant**, as it limits the duration and hence the amount of product that a reaction can produce
- The amount of product is therefore **directly proportional** to the amount of the limiting reactant added at the beginning of a reaction
- The limiting reactant is the reactant which is **not present in excess** in a reaction
- In order to determine which reactant is the limiting reactant in a reaction, we have to consider the ratios of each reactant in the balanced equation
- When performing reacting mass calculations, the limiting reactant is always the number that should be used as it indicates the maximum possible amount of product
- The steps are:
 1. Write the balanced equation for the reaction
 2. Calculate the moles of each reactant
 3. Compare the moles & deduce the limiting reactant


Your notes

Worked example

9.2 g of sodium is reacted with 8.0 g of sulfur to produce sodium sulfide, Na_2S.

Which reactant is in excess and which is the limiting reactant?

Relative atomic masses (A_r): Na = 23; S = 32

Answer

Step 1: Write the balanced equation and determine the molar ratio

$$2Na + S \rightarrow Na_2S \text{ so the molar ratio of Na : S is 2 : 1}$$

Step 2: Calculate the moles of each reactant

$$\text{Moles} = \text{Mass} \div \text{Molar Mass}$$

$$\text{Moles Na} = 9.2 \div 23 = 0.40$$

$$\text{Moles S} = 8.0 \div 32 = 0.25$$

Step 3: Compare the moles

- To react completely 0.40 moles of Na requires 0.20 moles of S and since there are 0.25 moles of S, then S is in excess
- **Na is therefore the limiting reactant**

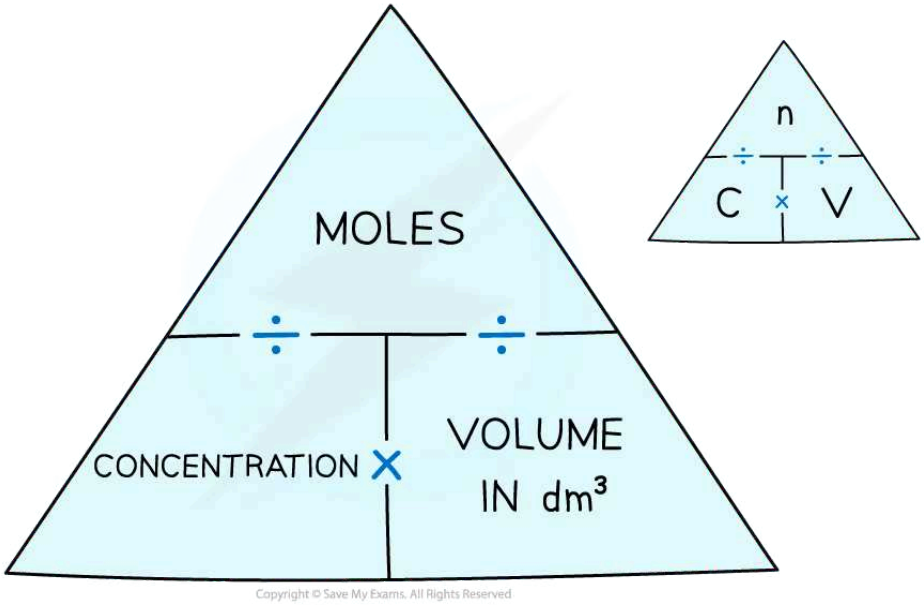
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Titration Calculations

Your notes

Titration Calculations

- Titrations are a method of analysing the **concentration** of solutions
- Acid-base titrations are one of the most important kinds of titrations
- They can determine exactly how much alkali is needed to neutralise a quantity of acid – and vice versa
- You may be asked to calculate the **moles** present in a given amount, the **concentration** or **volume** required to **neutralise** an acid or a base
- Once a titration is completed and the average titre has been calculated, you can now proceed to calculate the unknown variable using the formula triangle as shown below



Formula triangle showing the relationship between concentration, number of moles and volume of liquid

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Worked example

A solution of 25.0 cm^3 of hydrochloric acid was titrated against a solution of 0.100 mol/dm^3 NaOH and 12.1 cm^3 were required for complete reaction. Determine the concentration of the acid.

Answer

Step 1: Write the equation for the reaction:

$$\mathbf{HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)}$$

Step 2: Calculate the number of moles of the NaOH

Moles = (volume ÷ 1000) x concentration

Moles of NaOH = 0.012 dm^3 x 0.100 mol/dm^3 = 1.21 x 10^{-3} mol

Step 3: Deduce the number of moles of the acid

Since the acid reacts in a 1:1 ratio with the alkali, the number of moles of HCl is also 1.21 x 10^{-3} mol

This is present in 25.0 cm^3 of the solution (25.0 cm^3 = 0.025 dm^3)

Step 4: Find the concentration of the acid

Concentration = moles ÷ volume (dm^3)

Concentration of HCl =1.21 x 10^{-3} mol ÷ 0.025 dm^3 = **0.0484 mol/dm^3**

Your notes

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Empirical & Molecular Formula


Your notes

Calculating Empirical & Molecular Formulae

Calculating Empirical Formula

- The **empirical formula** is the simplest whole number ratio of the atoms of each element present in one molecule or formula unit of the compound
 - E.g. the empirical formula of ethanoic acid is CH_2O
- **Organic molecules** often have **different** empirical and molecular formulae
- The formula of an **ionic compound** is always an **empirical formula**

Worked example

A compound that contains 10 g of hydrogen and 80 g of oxygen.

What is its empirical formula of this compound?

Answer

	hydrogen	oxygen
Write the mass of each element	10 g	80 g
Divide each mass by the relative atomic mass to find the number of moles	10/1 = 10	80/16 = 5
Find the molar ratio by dividing by the smallest number	10/5 = 2	5/5 = 1

Empirical formula = H_2O

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Your notes

Worked example

Substance **X** was analysed and found to contain 31.58% carbon, 5.26% hydrogen and 63.16% oxygen by mass.

What is the empirical formula of substance **X**?

Relative atomic masses, Ar: C = 12; H = 1; O = 16

Answer

	carbon	hydrogen	oxygen
Convert % to g by assuming 100 g of substance is present	31.58 g	5.26 g	63.16 g
Divide each mass by the relative atomic mass to find the number of moles in 100 g	31.58/12 = 2.63	5.26/1 = 5.26	63.16/16 = 3.95
Find the molar ratio by dividing by the smallest number	2.63/2.63 = 1	5.26/2.63 = 2	3.95/2.63 = 1.5
Multiply all by 2 to obtain a whole number ratio	2	4	3

Empirical formula = $C_2H_4O_3$

Exam Tip

The molar ratio must be a whole number. If you don't get a whole number when calculating the ratio of atoms in an empirical formula, such as 1.5, multiply that and the other ratios to achieve whole numbers.

Calculating Molecular Formula

- Molecular formula gives the **actual numbers of atoms of each element** present in the formula of the compound
- To calculate the molecular formula:
 - Step 1: Find the relative formula mass of the empirical formula
 - Step 2: Use the following equation:

$$\frac{\text{relative formula mass of molecular formula}}{\text{relative formula mass of empirical formula}}$$

 - Step 3: Multiply the number of each element present in the empirical formula by the number from step 2 to find the molecular formula

Table showing the Relationship between Empirical and Molecular Formula

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Your notes

NAME OF COMPOUND	EMPIRICAL FORMULA	MOLECULAR FORMULA
METHANE	CH_4	CH_4
ETHANE	CH_3	C_2H_6
ETHENE	CH_2	C_2H_4
BENZENE	CH	C_6H_6

Worked example

The empirical formula of X is $C_4H_{10}S_1$

The relative formula mass (M_r) of X is 180.

What is the molecular formula of X?

(Relative atomic mass, A_r: Carbon : 12 Hydrogen : 1 Sulfur : 32)

Answer

Step 1 – Calculate the relative formula mass of the empirical formula

$$M_r = (12 \times 4) + (1 \times 10) + (32 \times 1) = 90$$

Step 2 – Divide relative formula mass of X by relative formula mass of empirical formula

$$180 / 90 = 2$$

Step 3 – Multiply each number of elements by 2

$$(C_{4 \times 2}) + (H_{10 \times 2}) + (S_{1 \times 2})$$

Molecular Formula of X = $C_8H_{20}S_2$

Deducing formulae of hydrated salts

- The formula of hydrated salts can be determined experimentally by weighing a sample of the hydrated salt, heating it until the water of crystallisation has been driven off, then reweighing the now anhydrous

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salt

- From the results, you can determine the mass of anhydrous salt and the mass of the water of crystallisation
- Applying a similar approach to deducing empirical formulae, the formula of the hydrated salt can be calculated

Your notes

Worked example

11.25 g of hydrated copper sulfate, $CuSO_4.xH_2O$, is heated until it loses all of its water of crystallisation. It is reweighed and its mass is 7.19 g. What is the formula of the hydrated copper(II) sulfate?

Answer

	$CuSO_4$	H_2O
Deduce the mass of water of crystallisation and anhydrous salt	= mass of salt after heating = 7.19 g	mass of hydrated salt - mass of anhydrous salt: 11.25 - 7.19 = 4.06 g
Divide each mass by the relative formula mass to find the number of moles	$\frac{7.19}{160} = 0.045$	$\frac{4.06}{18} = 0.226$
Find the molar ratio by dividing by the smallest number	$\frac{0.045}{0.045} = 1$	$\frac{0.226}{0.045} = 5$

Therefore the formula of hydrated copper(II) sulfate is $CuSO_4.5H_2O$

Exam Tip

The specification is not clear about whether deducing the formula of hydrated salts is required, however, it is an application of deducing empirical formulae so it is worth knowing how to do this.

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Percentage Yield & Purity

Your notes

Calculating Percentage Yield, Percentage by Mass & Percentage Purity

Percentage Yield

- **Yield** is the term used to describe the amount of **product** you get from a reaction
- In practice, you **never** get 100% yield in a chemical process for several reasons
- These include:
 - Some reactants may be left behind in the **equipment**
 - The reaction may be **reversible** and in these reactions a high yield is never possible as the products are continually turning back into the reactants
 - Some products may also be lost during **separation** and **purification** stages such as filtration or distillation
 - There may be **side reactions** occurring where a substance reacts with a gas in the air or an **impurity** in one of the reactants
 - Products can also be lost during **transfer** from one container to another

Actual & Theoretical Yield

- The **actual yield** is the recorded amount of product obtained
- The **theoretical yield** is the amount of product that would be obtained under perfect practical and chemical conditions
- It is calculated from the balanced equation and the reacting masses
- The **percentage yield** compares the actual yield to the theoretical yield
- For economic reasons, the objective of every chemical producing company is to have as high a percentage yield as possible to increase profits and reduce costs and waste
- To calculate percentage yield the following equation is used:

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

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Your notes

Worked example

Copper(II) sulfate may be prepared by the reaction of dilute sulfuric acid on copper(II) oxide.

A student prepared 1.6 g of dry copper(II) sulfate crystals.

Calculate the percentage yield if the theoretical yield is 2.0 g.

Answer

- Actual yield of copper(II) sulfate = 1.6 g
- Percentage yield of copper(II) sulfate = (1.6 / 2.0) x 100
- Percentage yield = **80%**

Exam Tip

Percentage yield cannot be greater than 100%, if you calculate a percentage higher than this, you have made an error! The most common error is to divide the theoretical yield by the actual yield so you may just need to swap the number around in your calculation.

Calculating percentage mass

- You may be asked to find the percentage by mass of an element within a compound
- This can be found by calculating the relative formula mass of the compound and then using the following equation:

$$\%\ \text{of an element} = \frac{\text{total mass of the element in the compound}}{\text{relative formula mass of the compound}} \times 100$$

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Your notes

Worked example

Calculate the percentage by mass of iron in iron(III) oxide, Fe_2O_3.

Relative atomic masses, A_r: Fe = 56 O = 16

Answer

Step 1: Calculate the relative formula mass of the compound

$$(2 \times 56) + (3 \times 16) = 160$$

Step 2: Work out the total A_r of the element you have been asked about in the question, in this case, iron

$$2 \times 56 = 112$$

Step 3: Divide the total A_r of the element by the M_r of the compound, then multiply by 100.

$$(112 \div 160) \times 100 = \mathbf{70\%}$$

Percentage Purity

- A pure substane has nothing else mixed with it
- Often, the product you are trying to obtain may become contaminated with unwanted substances such as unreacted reactants, catalysts and other impurities
- To calculate the percentage purity the following equation is used:

$$\%\ \text{purity} = \frac{\text{mass of pure substance}}{\text{total mass of substance}} \times 100$$

Worked example

A sample of lead(II) bromide was made. It weighed 15 g.

The sample was found to be impure and only contained 13.5 g of lead(II) bromide.

Calculate the percentage purity of the lead(II) bromide.

Answer

- The total mass of the substance is 15 g
- The mass of the pure substance is 13.5 g
- To calculate the percentage purity: (13.5 / 15) x 100
- Percentage purity= **90 %**

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Exam Tip

All of these calculations are to find a percentage so don't forget to multiply by 100 to convert your answer to a percentage.



Your notes

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Cambridge O Level Chemistry

Your notes

4.1 Electrolysis

Contents

- Electrolysis Principles
- Electrolysis of Molten Compounds
- Electrolysis of Aqueous Sodium Chloride & Dilute Sulfuric Acid
- Electrolysis of Aqueous Solutions
- Ionic Half Equations
- Electroplating

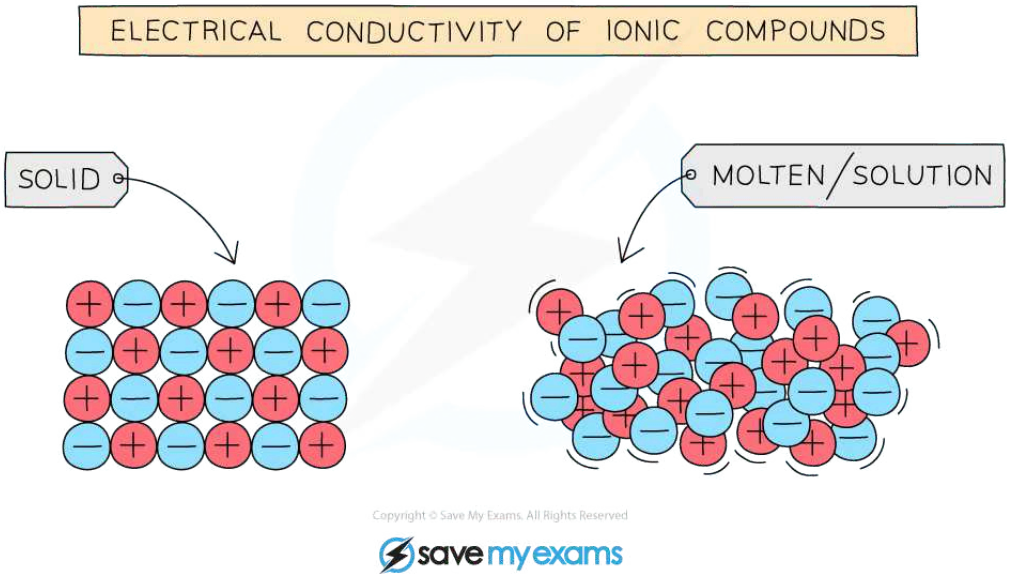
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Electrolysis Principles

Your notes

Electrolysis: General Principles

- When an electric current is passed through a **molten ionic compound** the compound decomposes or breaks down
- The process also occurs for **aqueous solutions** of ionic compounds
- Covalent compounds cannot conduct electricity hence they do not undergo electrolysis
- Ionic compounds in the solid state cannot conduct electricity either since they have **no free ions** that can move and carry the charge



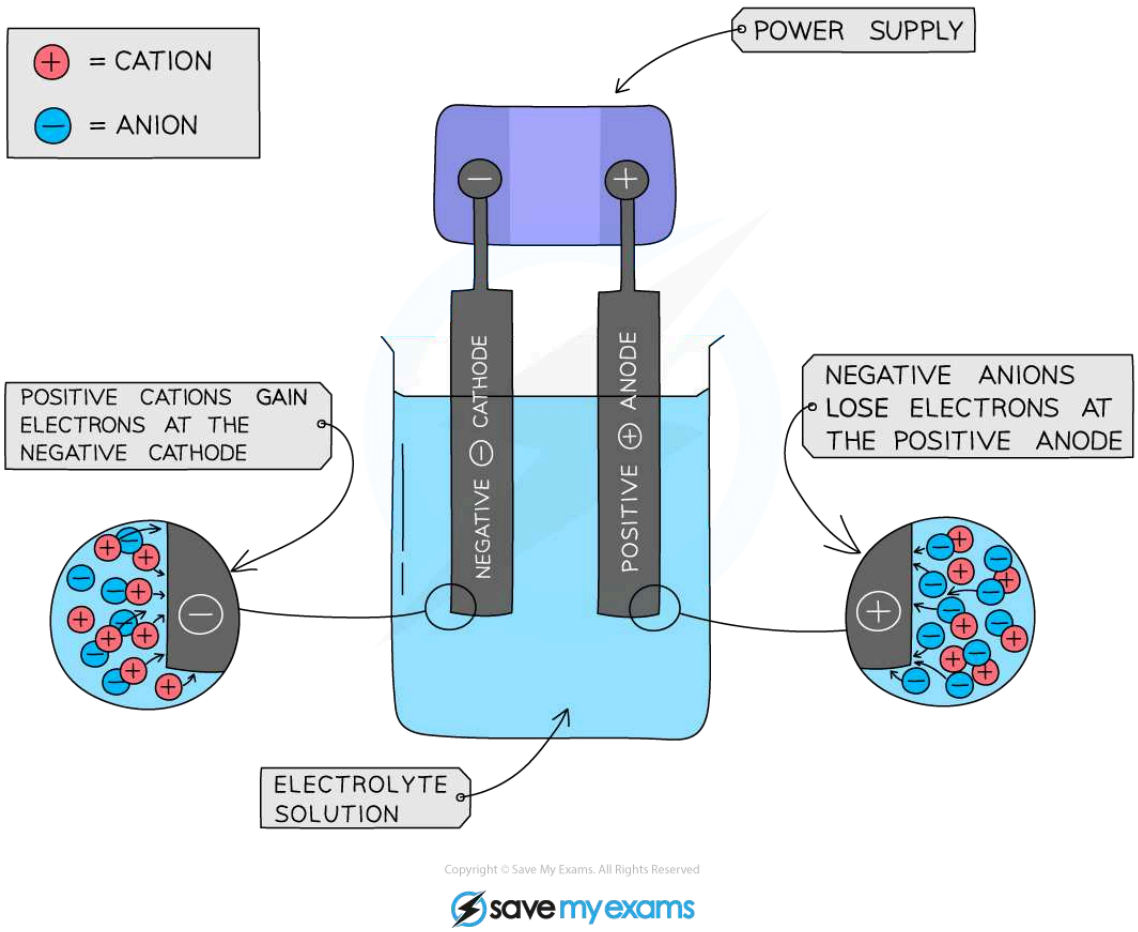
Particles in ionic compounds are in fixed position in the solid state but can move around when molten or in solution

Key terms used in a simple electrolytic cell

- **Electrode** is a rod of metal or graphite through which an electric current flows into or out of an electrolyte
- **Electrolyte** is the ionic compound in a molten or dissolved solution that conducts the electricity
- **Anode** is the positive electrode of an electrolysis cell
- **Anion** is a negatively charged ion which is attracted to the anode
- **Cathode** is the negative electrode of an electrolysis cell
- **Cation** is a positively charged ion which is attracted to the cathode

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The basic set-up of an electrolytic cell

- Metals and hydrogen form positively charged ions and so either a **metal** or **hydrogen** gas is formed at the **cathode**
- Non-metals form negatively charged ions and so **non-metals (except hydrogen)** are formed at the **anode**

Exam Tip

Use the PANIC mnemonic to remember which electrode is the positive and which is the negative:

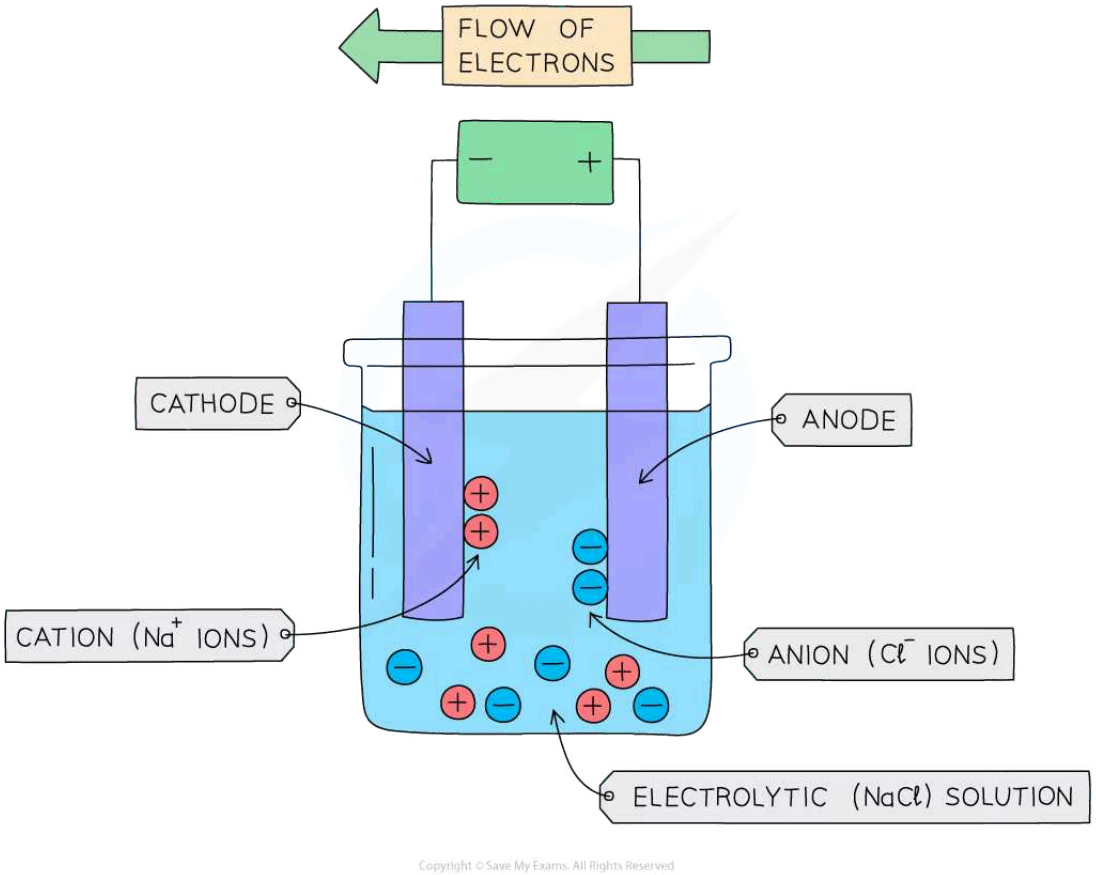
Positive (is) Anode Negative Is Cathode

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Electrolysis: Charge Transfer

Your notes

- During electrolysis, current needs to flow around the circuit
- In order for this to occur, charge must be transferred around the circuit (current is a measure of the rate of flow of charge) by charge carriers
- The power supply provides the **cathode** with a supply of **electrons**, causing it to become negatively charged
- **Positive ions** (cations) in the electrolyte move towards the cathode where they **gain electrons**
- **Negative ions** (anions) in the electrolyte move towards the anode where they **lose electron**s
- The **electrons** move from the anode back towards the power supply
- So, in a complete circuit:
 - **Electrons** are the charge carriers in the **external circuit**
 - **Ions** are the charge carriers in the **electrolyte**



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Diagram showing the direction of movement of electrons and ions in the electrolysis of NaCl

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Electrolysis of Molten Compounds


Your notes

Electrolysis of Molten Compounds

- A binary ionic compound is one consisting of just two elements joined together by ionic bonding
- When these compounds undergo electrolysis they always produce their corresponding elements
- To predict the products made at each electrode, first identify the ions
- The **positive** ion will migrate towards the **cathode** and the **negative** ion will migrate towards the **anode**
- Therefore, the **cathode** product will always be the **metal,** and the product formed at the **anode** will always be the **non-metal**

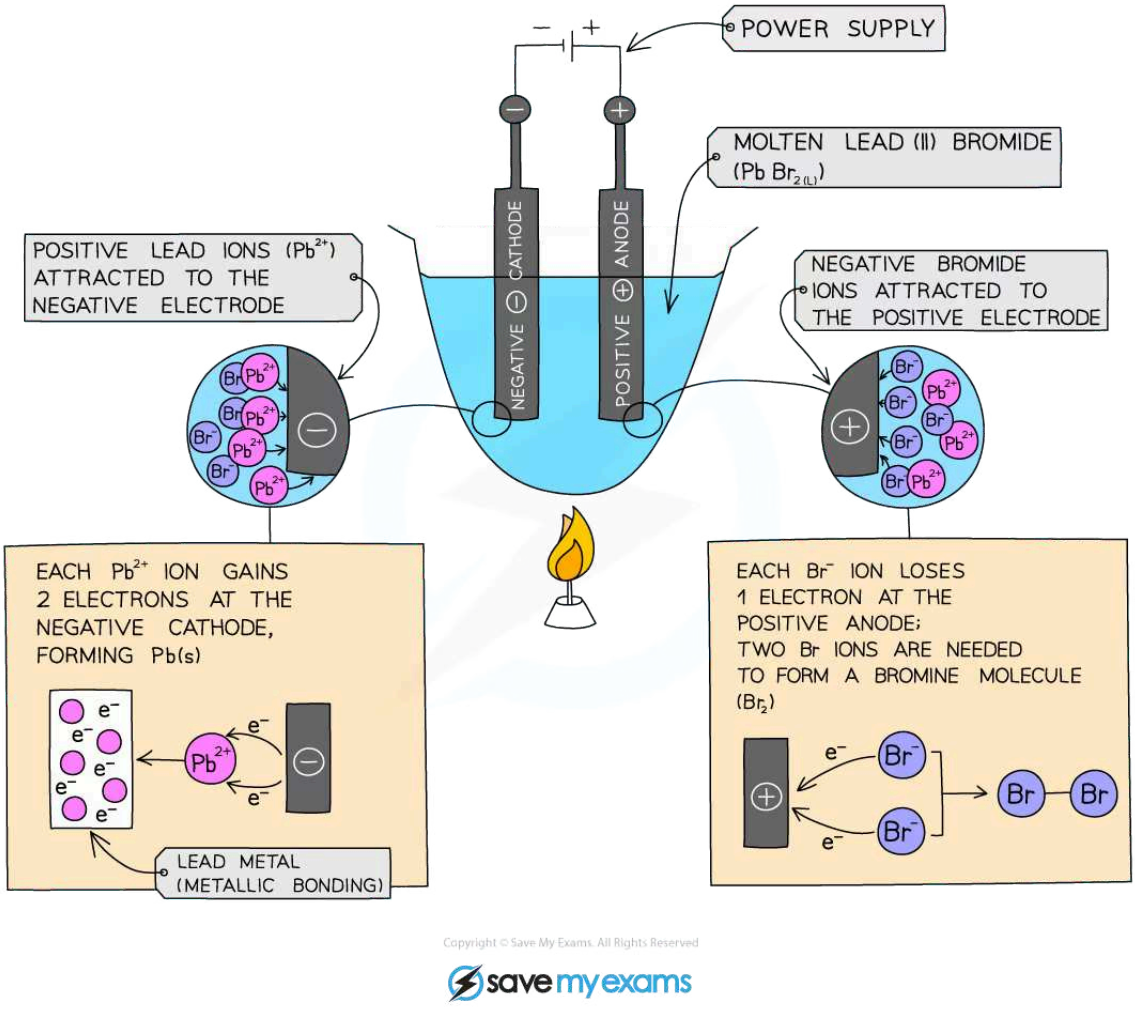
Example: Electrolysis of molten lead(II) bromide

Method:

- Add lead(II) bromide into a beaker and heat it so it will turn molten, allowing ions to be free to move and conduct an electric charge
- Add two graphite rods as the electrodes and connect this to a power pack or battery
- Turn on the power pack or battery and allow electrolysis to take place
- Negative bromide ions move to the positive electrode (anode) and each loses one electron to form bromine molecules. There is bubbling at the anode as brown **bromine gas is given off**
- Positive lead ions move to the negative electrode (cathode) and gain electrons to form a **grey lead metal** which deposits on the surface of the electrode

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Diagram showing the electrolysis of lead(II) bromide

Worked example

Identify the product formed at the anode and cathode during the electrolysis of molten potassium chloride.

Answer

- The ions present are potassium (K^+) and chloride (Cl^-)
- The chloride ions are attracted to the anode and form chlorine gas
- The potassium ions are attracted to the cathode and form potassium metal

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Exam Tip

Remember electrodes need to be inert such as graphite or platinum so that they don't participate in a side reaction with the electrolyte.



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Electrolysis of Aqueous Sodium Chloride & Dilute Sulfuric Acid

Your notes

Electrolysis of Aqueous Sodium Chloride & Dilute Sulfuric Acid

Aqueous sodium chloride

- **Brine** is a concentrated solution of aqueous sodium chloride
- It can be electrolysed using inert electrodes made from platinum or carbon/graphite
- When electrolysed, it produces bubbles of gas at both electrodes as chlorine and hydrogen are produced, leaving behind sodium hydroxide solution
- These substances all have important industrial uses:
 - Chlorine is used to make bleach
 - Hydrogen is used to make margarine
 - Sodium hydroxide is used to make soap and detergents

Product at the Negative Electrode:

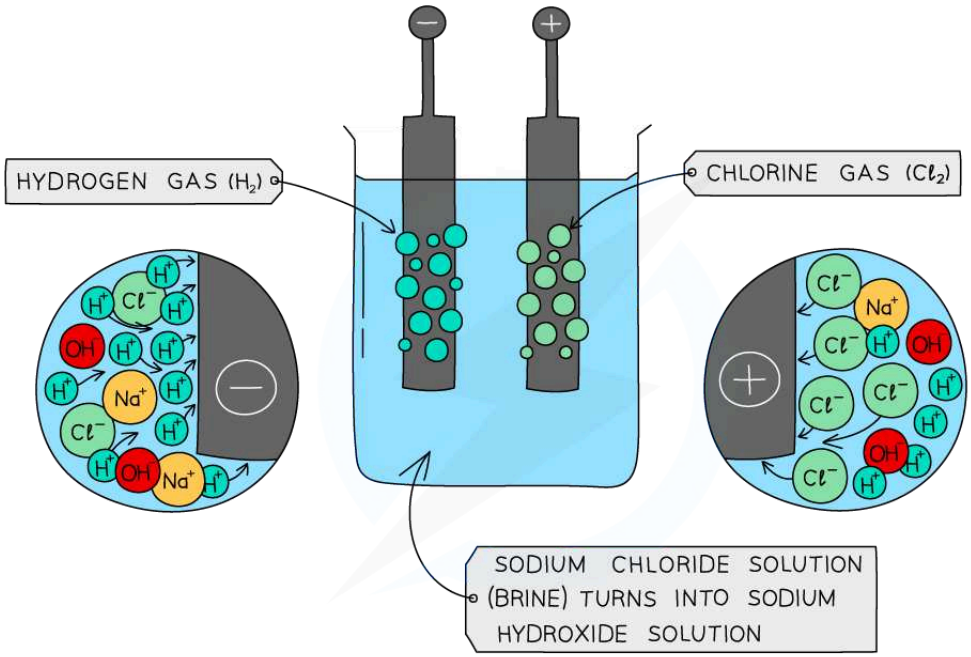
- The H^+ ions are discharged at the **cathode** as they are less reactive than sodium ions
- The H^+ ions gain electrons to form hydrogen gas

Product at the Positive Electrode:

- The Cl^- ions are discharged at the **anode**
- They lose electrons and **chlorine gas forms**
- The Na^+ and OH^- ions remain behind and form the NaOH solution

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Diagram showing the products of the electrolysis of aqueous sodium chloride

Dilute sulfuric acid

- Dilute sulfuric acid can be electrolysed using inert electrodes made from platinum or carbon/graphite
- Bubbles of gas are seen at both electrodes

Product at the Negative Electrode

- H^+ ions are attracted to the **cathode,** gain electrons and form **hydrogen gas**

Product at the Positive Electrode

- OH^- ions are attracted to the **anode**, lose electrons and form **oxygen gas** and water

Determining what gas is produced

- If the gas produced at the **anode** relights a glowing splint dipped into a sample of the gas then the gas is **oxygen**

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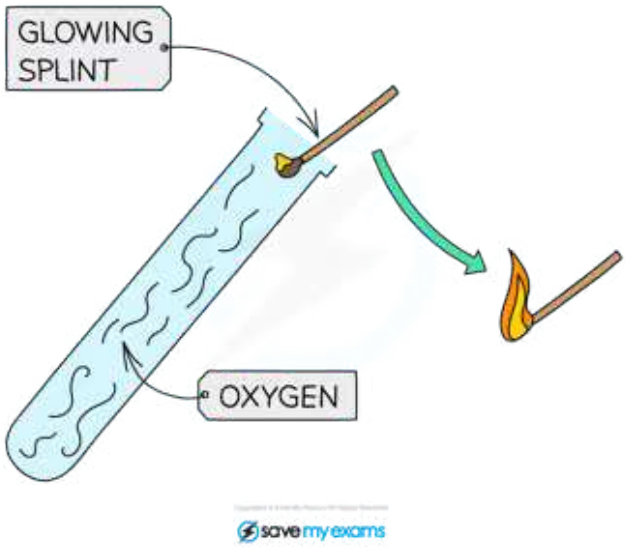


Diagram showing the test for oxygen gas

- If the gas produced at the **anode** bleaches damp litmus paper then the gas is **chlorine**

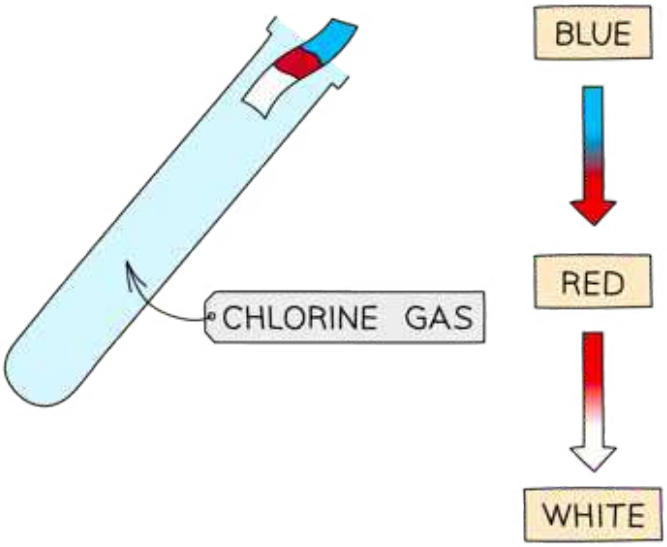


Diagram showing the test for chlorine gas

- If the gas produced at the **cathode** burns with a 'pop' when a sample is lit with a lighted splint then the gas is **hydrogen**

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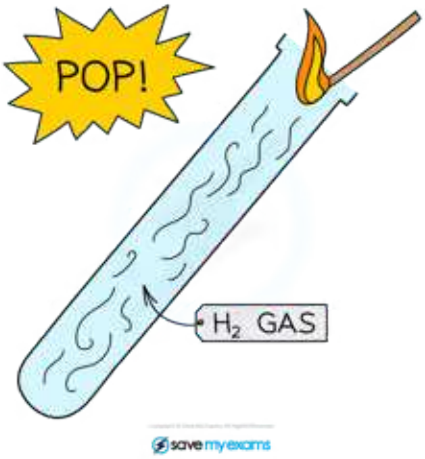


Diagram showing the test for hydrogen gas

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Electrolysis of Aqueous Solutions

Your notes

Electrolysis of Aqueous Solutions

- Aqueous solutions will always have water present (H_2O)
- In the electrolysis of aqueous solutions, the water molecules dissociate producing H^+ and OH^- ions:

$$H_2O \rightleftharpoons H^+ + OH^-$$

- These ions are also involved in the process and their chemistry must be considered
- We now have an electrolyte that contains ions from the compound plus ions from the water
- Which ions get discharged and at which electrode depends on the **relative reactivity** of the elements involved
- Concentrated and dilute solutions of the **same** compound give **different** products
- For anions, the **more concentrated** ion will tend to get discharged over a more dilute ion

Positive Electrode (anode)

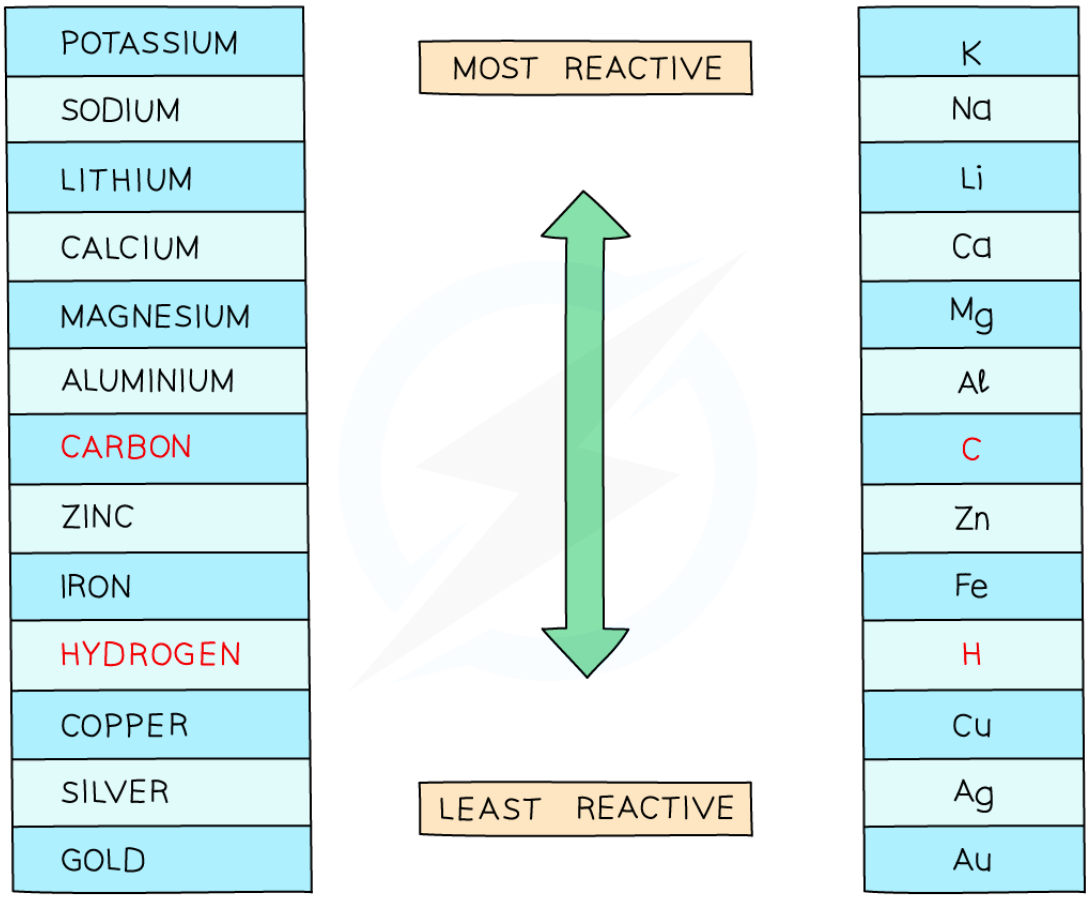
- Negatively charged OH^- ions and non-metal ions are attracted to the positive electrode
- If halide ions (Cl^-, Br^-, I^-) and OH^- are present then the halide ion is discharged at the anode, loses electrons and forms a halogen (chlorine, bromine or iodine)
- If no halide ions are present, then OH^- is discharged at the anode, loses electrons and forms oxygen gas
- In both cases the other negative ion remains in solution
- The concentration of the solution also affects which ion is discharged:
 - If a **concentrated** halide solution is being electrolysed, the **halogen** forms at the anode
 - If a **dilute** halide solution is being electrolysed, **oxygen** is formed
- For example:
 - For a concentrated solution of barium chloride, the Cl^- ions are discharged more readily than the OH^- ions, so chlorine gas is produced at the anode
 - If the solution is dilute however only the OH^- ion is discharged and so oxygen would be formed

Negative Electrode (cathode)

- Positively charged H^+ and metal ions are attracted to the negative electrode but only one will gain electrons
- Either hydrogen gas or metal will be produced
- If the metal is **above hydrogen** in the reactivity series, then hydrogen will be produced and bubbling will be seen at the cathode
- This is because the ions of the more reactive metal will remain in the solution, causing the ions of the least reactive metal to be discharged
- Therefore, at the cathode, **hydrogen gas** will be produced unless the positive ions from the ionic compound are less reactive than hydrogen, in which case the **metal** is produced

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The reactivity series of metals including hydrogen and carbon

Products formed for Common Aqueous Solutions

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Aqueous Solution (Ions present)	Product at Anode	Product of Cathode
Concentrated Sodium Chloride (NaCl)	Chlorine gas released	Hydrogen gas released
Dilute Sodium Chloride (NaCl)	Oxygen produced	Hydrogen gas released
Concentrated aqueous Copper(II) Sulfate ($CuSO_4$)	Oxygen gas released	Copper is lower than hydrogen in the reactivity series so copper is preferentially discharged as a metal
Dilute Sulfuric Acid (H_2SO_4)	Oxygen gas released. H_2O more readily gives up electrons than SO_4^{2-}	Hydrogen gas released

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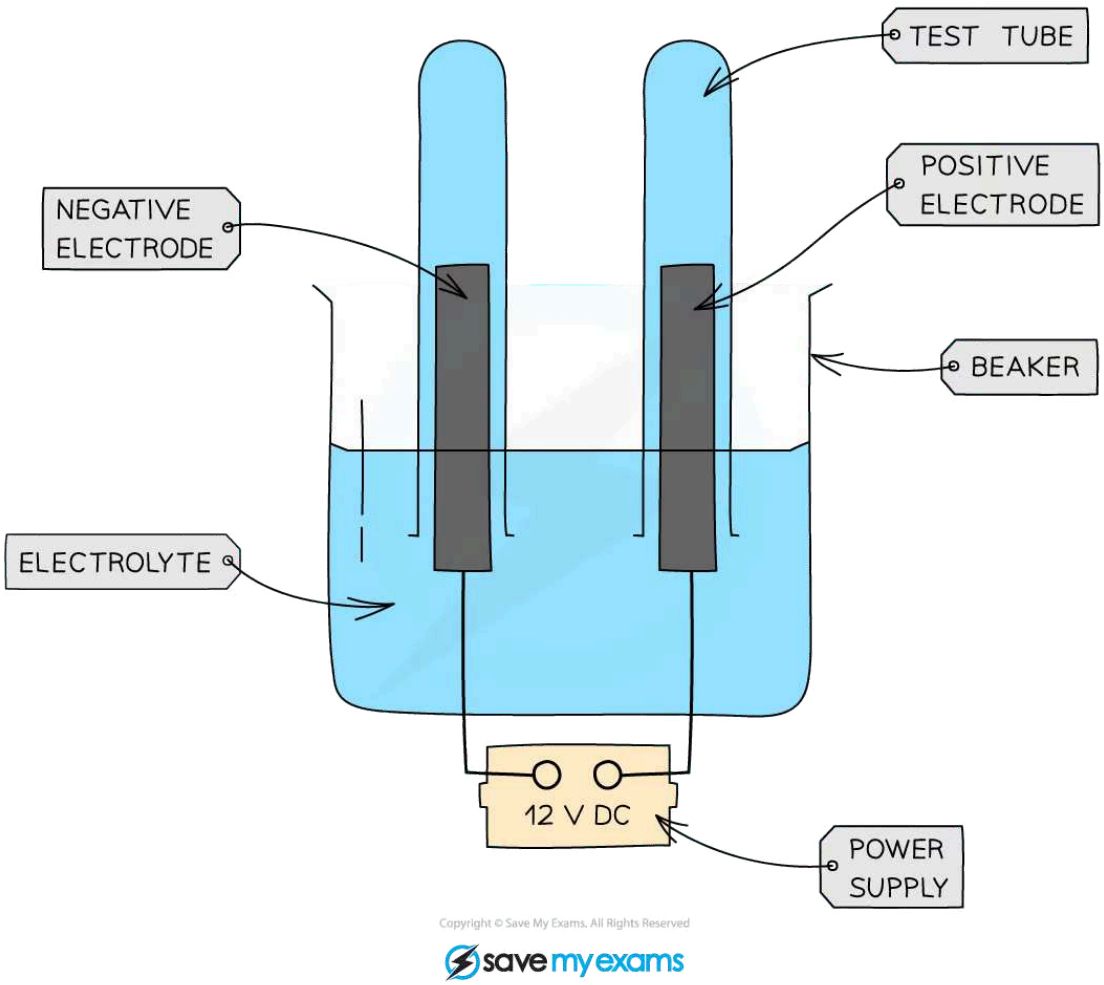
Electrolysis of Aqueous Copper(II) Sulfate

Your notes

Electrolysis of Aqueous Copper Sulfate

- Aqueous copper sulfate contains the following ions:
 - Cu^{2+}, SO_4^{2-}, H^+ and OH^-

Using graphite electrodes:



Apparatus for the electrolysis of copper(II)sulfate using inert / passive graphite electrodes

Product at the Cathode:

- Cu^{2+} and H^+ will both be attracted to the cathode but the less reactive ion will be discharged

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- In this case, copper is less reactive than hydrogen
- Copper ions are discharged at the cathode, gain electrons and are reduced to form copper metal
- The half equation for the reaction at the electrode is:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

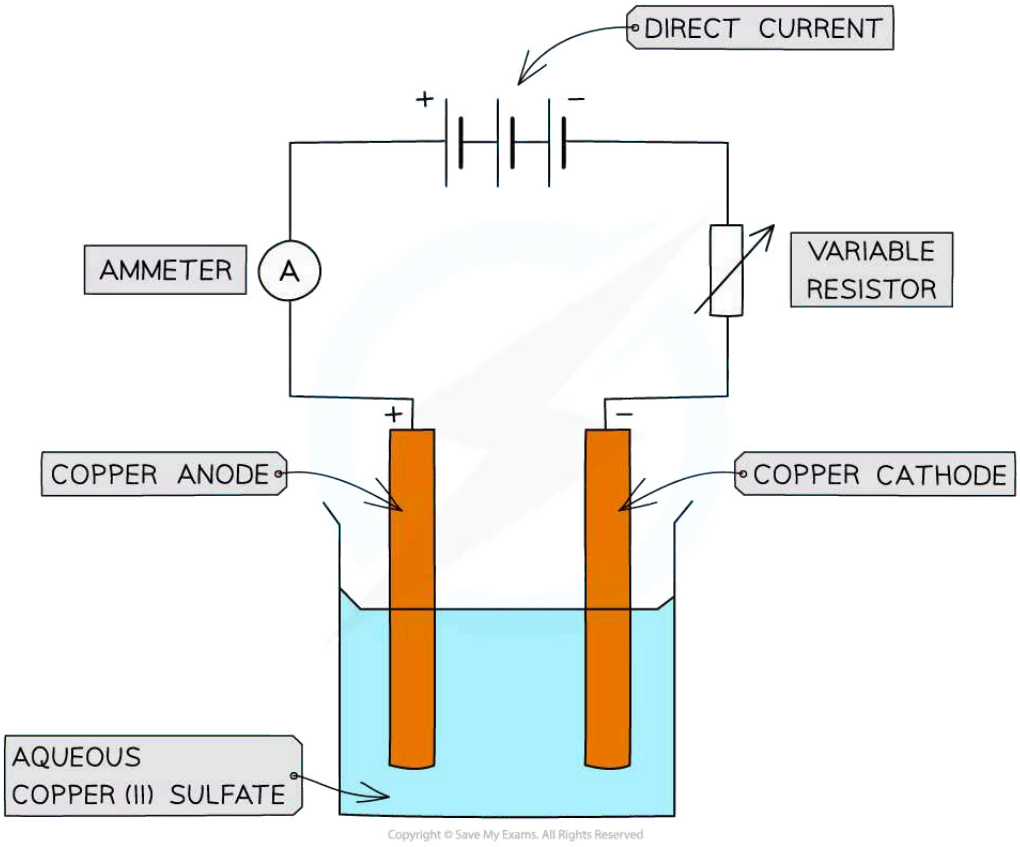
Your notes

Product at the Anode:

- SO_4^{2-} and OH^- are both attracted to the anode
 - OH- ions lose electrons more readily than SO_4^{2-}
 - OH- lose electrons and are oxidised to form oxygen gas
 - The half equation for the reaction at the anode is

$$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$$

Using copper electrodes:



Apparatus for the electrolysis of copper(II)sulfate using active copper electrodes

Observations at the anode and cathode

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- The **cathode increases** in mass while the **anode** decreases
- This occurs as copper atoms are **oxidised** at the anode and form ions while copper ions are **reduced** at the cathode, forming copper atoms
- The gain in mass by the negative electrode is the **same** as the loss in mass by the positive electrode
- Therefore the copper deposited on the negative electrode must be the **same** copper ions that are lost from the positive electrode
- That implies that the concentration of the Cu^{2+} ions in the solution remains **constant**


Your notes

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Ionic Half Equations

Your notes

Ionic Half Equations

- In electrochemistry we are mostly concerned with the transfer of electrons, hence the definitions of oxidation and reduction are applied in terms of electron loss or gain rather than the addition or removal of oxygen
- **Oxidation** is when a substance **loses electrons** and **reduction** is when a substance **gains electrons**
- As the ions come into contact with the electrode, electrons are either lost or gained and they form **neutral s**ubstances
- These are then discharged as products at the electrodes
- At the **anode**, negatively charged ions **lose electrons** and are thus oxidised
- At the **cathode,** the positively charged ions **gain electrons** and are thus reduced
- **Ionic half equations** show the oxidation and reduction of the ions involved
- It is important to make sure the charges are balanced

Table of Reduction and Oxidation Reactions at the Electrodes

	Anode Reaction	Cathode Reaction
Molten Lead Bromide	$2Br^- \longrightarrow Br_2 + 2e^-$	$Pb^{2+} + 2e^- \longrightarrow Pb$
Concentrated Aqueous Sodium Chloride	$2Cl^- \longrightarrow Cl_2 + 2e^-$	$2H^+ + 2e^- \longrightarrow H_2$
Dilute Sulfuric Acid	$4OH^- \longrightarrow O_2 + 2H_2O + 4e^-$	$2H^+ + 2e^- \longrightarrow H_2$

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Exam Tip

To help you remember the definitions of oxidation and reduction use OIL RIG

Oxidation **I**s **L**oss (of electrons) **R**eduction **I**s **G**ain (of electrons)

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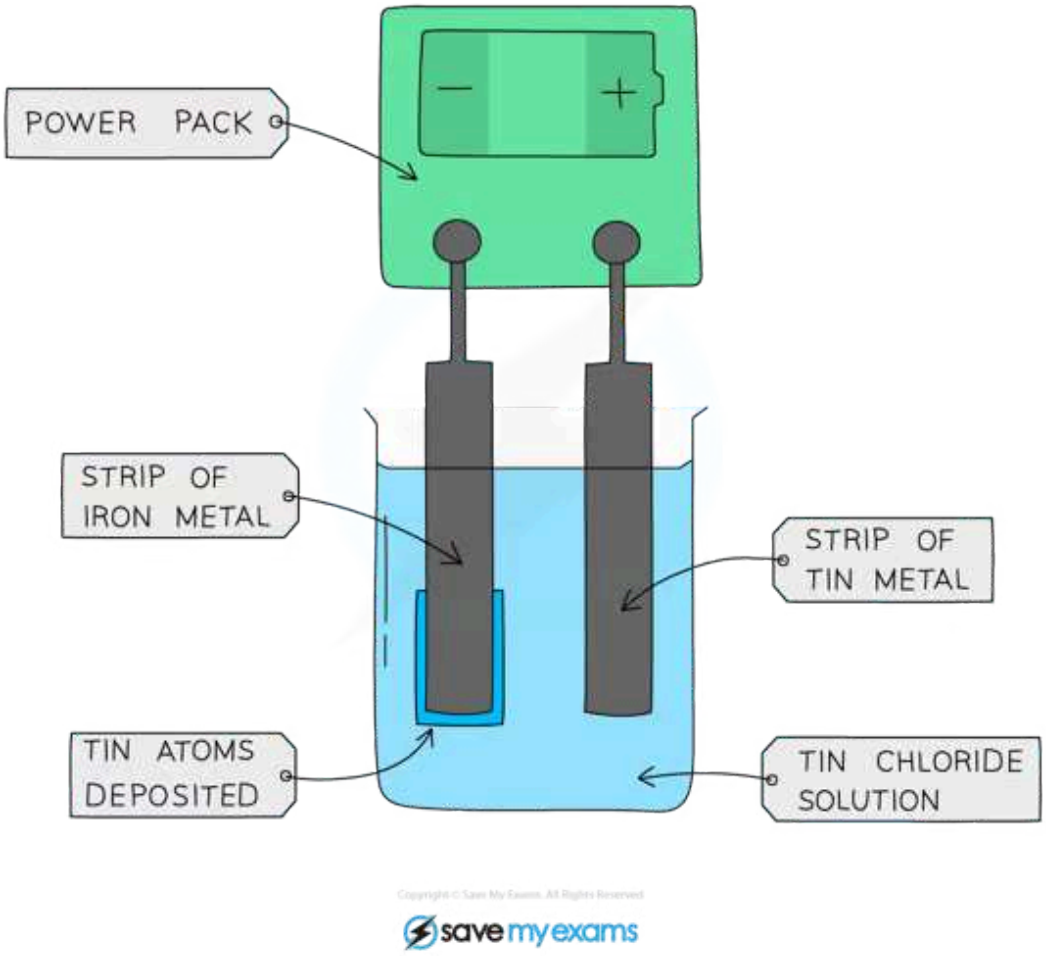
Electroplating

Your notes

Electroplating

- Electroplating is a process where the surface of one metal is **coated** with a layer of a different metal
- The **anode** is made from the **pure** metal you want to coat your object with
- The **cathode** is the **object** to be electroplated
- The electrolyte is an aqueous solution of a soluble salt of the pure metal at the anode

- **Example:** coating a strip of iron metal with tin:



A piece of iron being electroplated with tin. The electrolyte is tin(II) chloride, a water-soluble salt of tin

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- **At the anode:** Tin atoms lose electrons to form tin ions in solution
- **At the cathode:** Tin ions gain electrons to form tin atoms which deposit on the strip of iron metal, coating it with a layer of tin

Uses of electroplating

- Electroplating is done to make metals more **resistant** to corrosion or damage
 - e.g, chromium and nickel plating
- It is also done to **improve the appearance** of metals,
 - e.g. coating cutlery and jewellery with silver

Exam Tip

You may be asked to write the ionic half equations for the reaction at each electrode. For the example above, these would be:

At the anode: $Sn\,(s) \rightarrow Sn^{2+}\,(aq) + 2e^-$

At the cathode: $Sn^{2+}\,(aq) + 2e^- \rightarrow Sn\,(s)$


Your notes

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Cambridge O Level Chemistry

Your notes

4.2 Hydrogen Fuel Cells

Contents

Hydrogen Fuel Cells

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Hydrogen Fuel Cells


Your notes

Hydrogen Fuel Cells

- A fuel is a substance which releases energy when burned
- Hydrogen is used as a fuel in rocket engines and in fuel cells to power some cars
- A fuel cell is an electrochemical cell in which a fuel **donates** electrons at one electrode and oxygen **gains** electrons at the other electrode
 - $H_2 \rightarrow 2H^+ + 2e^-$
 - $O_2 + 4e^- \rightarrow 2O^{2-}$
- The **hydrogen-oxygen** fuel cell produces electricity by combining both elements, releasing energy and water
- The overall equation for the reaction within a hydrogen fuel cell is:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

- The diagram below shows the setup of a hydrogen fuel cell
- The air entering provides the oxygen
- The fuel entering is hydrogen
- The only chemical product made is **water**

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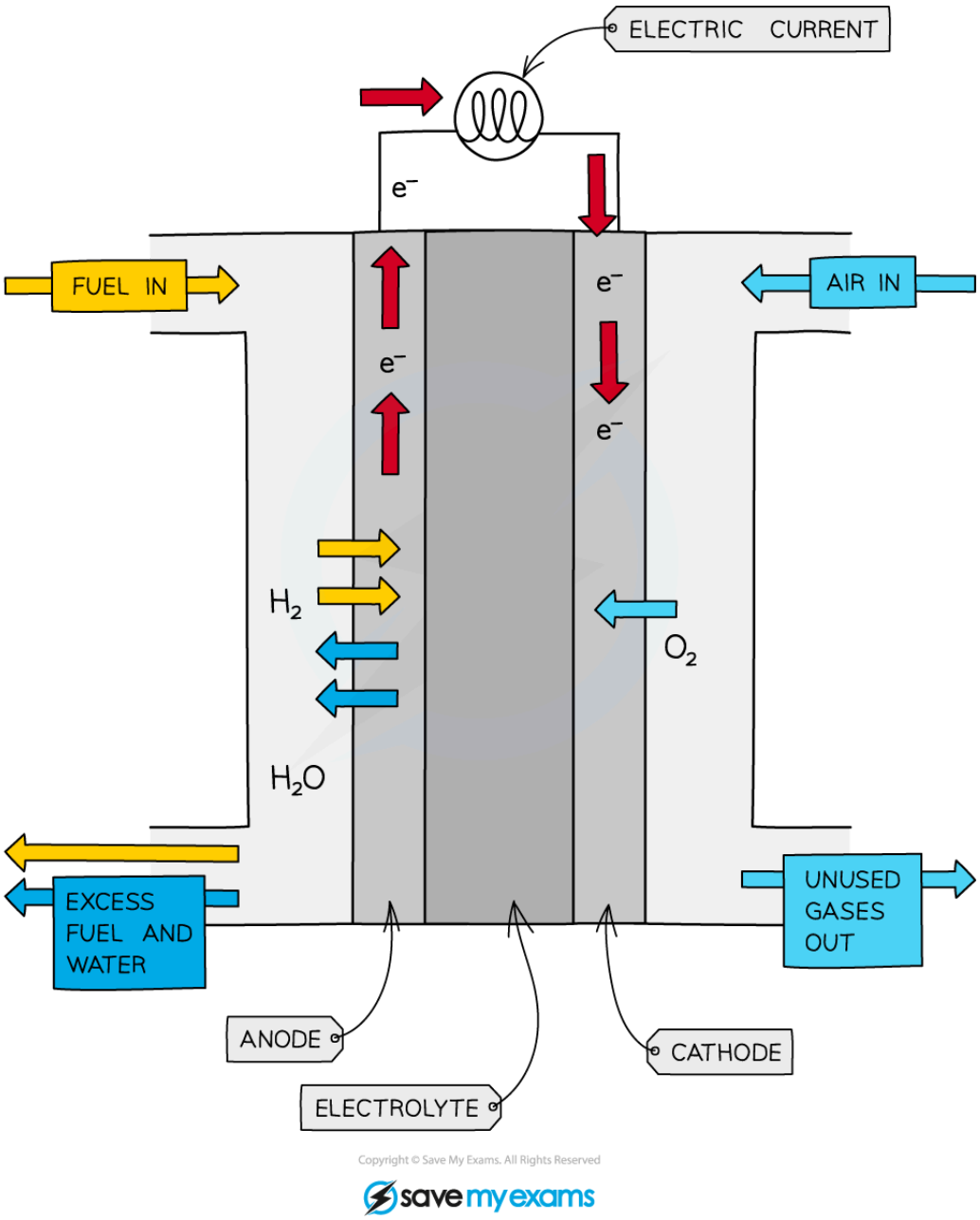


Diagram showing the movement of hydrogen, oxygen and electrons in a hydrogen-oxygen fuel cell

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Advantages & Disadvantages of Hydrogen Fuel Cells


Your notes

- Hydrogen-oxygen fuel cells are becoming more common in the automotive industry to replace petrol or diesel engines

Advantages

- They **do not produce any pollution**: the only product is water whereas petrol engines produce carbon dioxide, and oxides of nitrogen
- They release **more energy** per kilogram than either petrol or diesel
- No power is lost in transmission as there are no moving parts, unlike an internal combustion engine
- Quieter so less noise pollution compared to a petrol engine

Disadvantages

- Materials used in producing fuel cells are **expensive**
- Hydrogen is more difficult and expensive to store compared to petrol as it is very flammable and easily explodes when under pressure
- Fuel cells are affected by **low** temperatures, becoming less efficient
- There are only a small number of hydrogen filling stations across the country
- Hydrogen is often obtained by methods that involve the combustion of fossil fuels, therefore releasing carbon dioxide and other pollutants into the atmosphere

Exam Tip

You should be able to state advantages and disadvantages of the hydrogen-oxygen fuel cells in comparison to a petrol engine.

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Cambridge O Level Chemistry

Your notes

5.1 Exothermic & Endothermic Reactions

Contents

- Endothermic & Exothermic Reactions
- Enthalpy Change & Activation Energy
- Bond Breaking & Bond Forming

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Endothermic & Exothermic Reactions


Your notes

Exothermic & Endothermic Reactions

Heat Exchange in Reactions

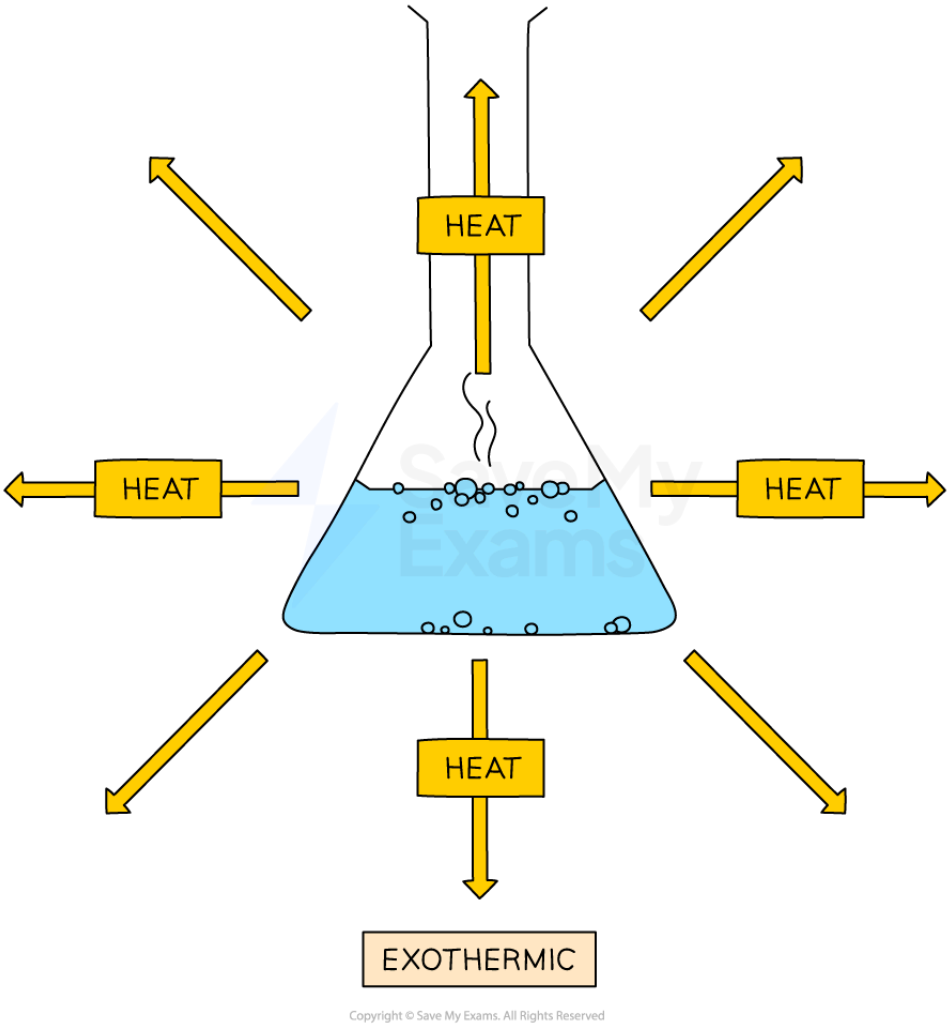
- Chemical reactions occur so that elements can achieve a more stable energy state by gaining a full outer shell of electrons
- This is done by **chemical bonding** (we have already seen ionic and covalent bonding) where old bonds are broken, and new bonds are formed)
- This process involves the **transfer of thermal energy** into and out of reaction mixtures
- The terms used to describe this are the **system** (what happens in the chemical reaction) and the **surroundings** (anything other than the chemicals reacting)
- The energy within the system comes from the **chemical bonds** themselves which could be considered as tiny stores of chemical energy

Exothermic Reactions

- In exothermic reactions thermal energy is **transferred to** the surroundings so the temperature of the surroundings increases
- This energy is transferred from the chemical energy store of the chemical system to the surroundings and so the energy of the system falls - this means that the energy change is negative
- The overall transfer is from the system to the surroundings
- Combustion, oxidation, and neutralisation reactions are typical exothermic reactions
- Hand warmers used in the wintertime are based on the release of heat from an exothermic reaction
- Self-heating cans of food and drinks such as coffee and hot chocolate also use exothermic reactions in the bases of the containers

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Diagram showing the transfer of heat energy outwards from an exothermic reaction

Endothermic Reactions

- In endothermic reactions thermal energy is **taken in from** the surroundings so the temperature of the surroundings decreases
- This energy is transferred **to** the chemical energy store of the chemical system and so the energy of the system increases - this means the energy change is positive
- The overall transfer is from the surroundings to the system
- These types of reactions are much less common than the exothermic reactions
- Electrolysis, thermal decomposition reactions and the first stages of photosynthesis are typical endothermic reactions

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- Sports injury treatments often use cold packs based on endothermic reactions to take heat away from a recently injured area to prevent swelling

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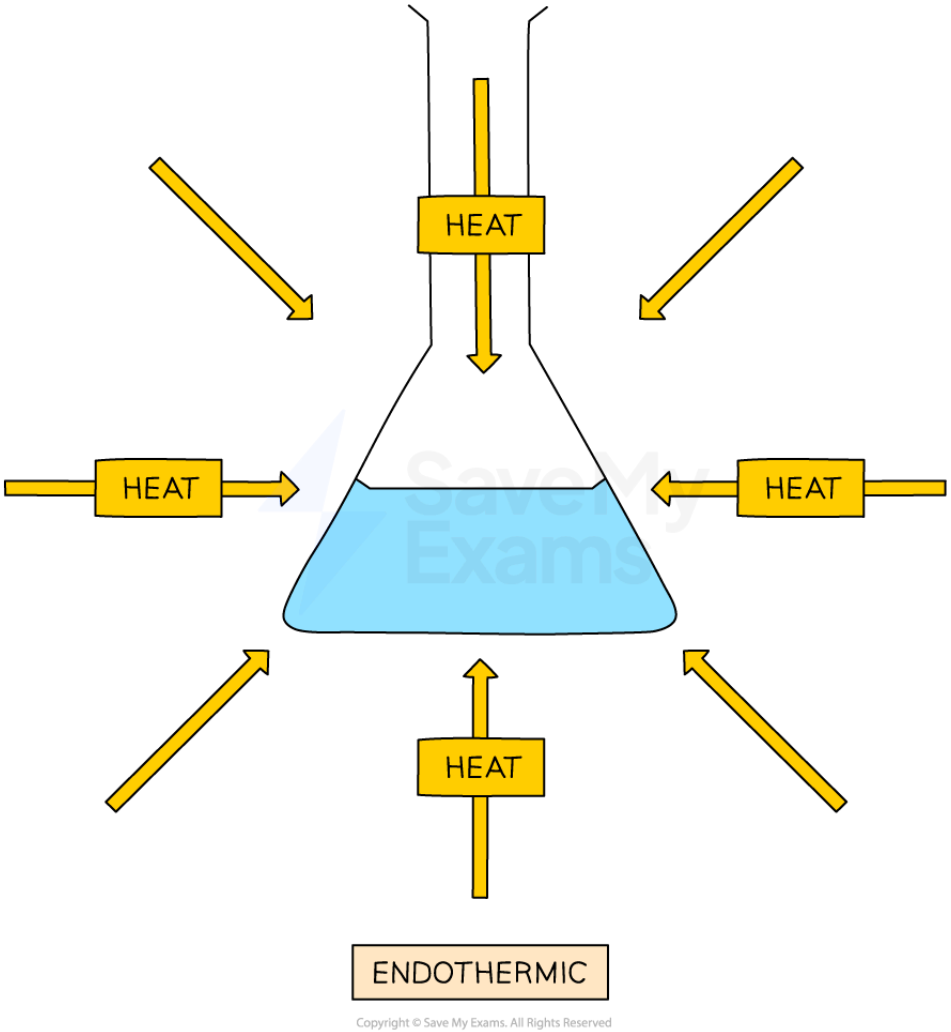


Diagram showing the transfer of heat energy from the surroundings into an endothermic reaction

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Exam Tip

To help you remember whether a chemical system is exothermic or endothermic:

- In **EX**othermic reactions heat **EX**its the system and in **EN**dothermic reactions heat **EN**ters the system.
- Exothermic reactions always give off heat and they feel hot, whereas endothermic reactions take heat in and they feel cold.



Your notes

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Enthalpy Change & Activation Energy


Your notes

Enthalpy Change & Activation Energy

- For atoms or particles to react with each other in a chemical system they must first of all come into contact with each other in a **collision**
- A number of factors come into play when analysing collisions such as **energy, orientation**, and **number of collisions per second** (the **frequency** of collisions)
- There is a **minimum amount of energy** that the particles must collide with for the collision to be successful, that is for the particles to react together
 - This minimum amount of energy is called the **activation energy (E_a)**
- Different reactions have different activation energies, depending on the chemical identities involved
- Reactions which have higher activation energies require more energy to start than those with lower activation energies
- The **transfer of thermal energy** during a reaction is called the **enthalpy change, ΔH**, of the reaction.
- ΔH is a positive value for endothermic reactions and a negative value for exothermic reactions

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Reaction Pathway Diagrams

Your notes

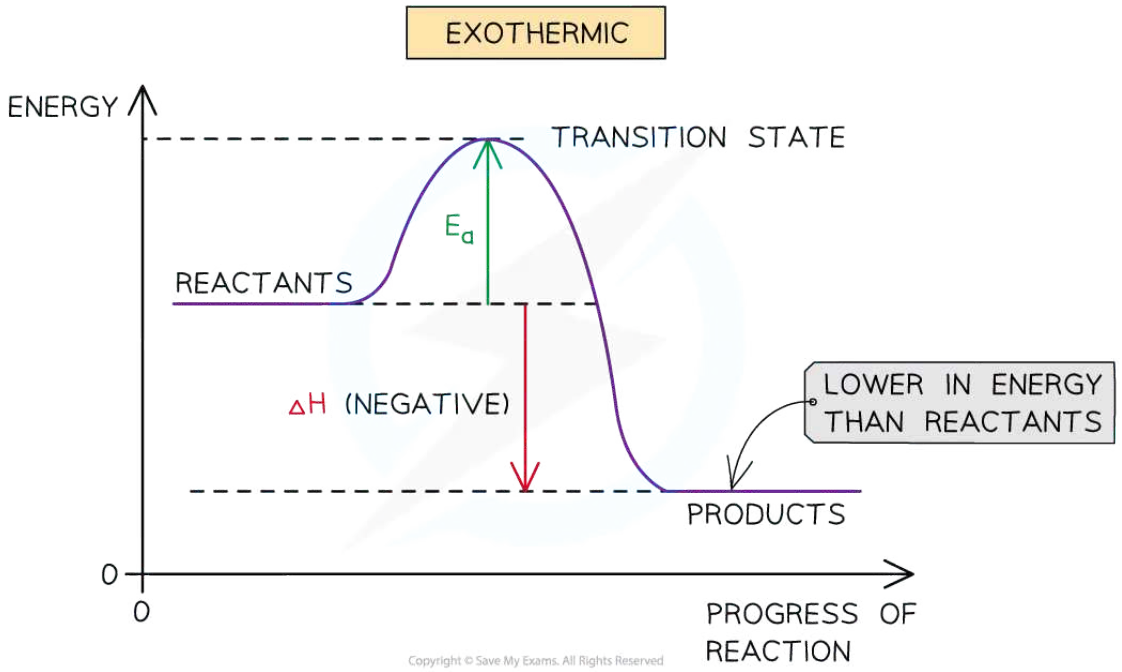
- **Reaction pathway diagrams** (a.k.a energy level diagrams) are graphical representations of the relative energies of the reactants and products in chemical reactions
- The energy of the reactants and products are displayed on the y-axis and the reaction pathway (a bit like time) is shown on the x-axis
- The difference in height between the energy of reactants and products represents the **overall energy change** of a reaction.
 - This is usually a sketch but can be drawn to scale if data is provided
- Arrows on the diagrams indicate whether the reaction is exothermic (overall reaction arrow is downwards pointing, showing that the system has lost energy) or endothermic (overall reaction arrow is upwards pointing, showing that the system has gained energy)
- The initial increase in energy represents the **activation energy (Ea)**, which is the minimum energy that colliding particles must have in order to react
- The greater the initial rise, the more energy that is required to get the reaction going e.g. more heat needed
- You can identify a reaction pathway diagram for an **exothermic reaction** as the energy of the **product** is **lower** than the reactants (as thermal energy has been transferred to the surroundings)
- You can identify a reaction pathway diagram for an **endothermic reaction** as the energy of the **product** is **higher** than the reactants (as thermal energy has been taken in from the surroundings)

Exothermic reactions

- If more energy is released than is absorbed, then the reaction is **exothermic**
- More energy is released when new bonds are formed than energy required to break the bonds in the reactants
- The change in energy is negative since the products have less energy than the reactants
- Therefore an exothermic reaction has a **negative** ΔH value
- The reaction pathway diagram for an exothermic reaction is shown below

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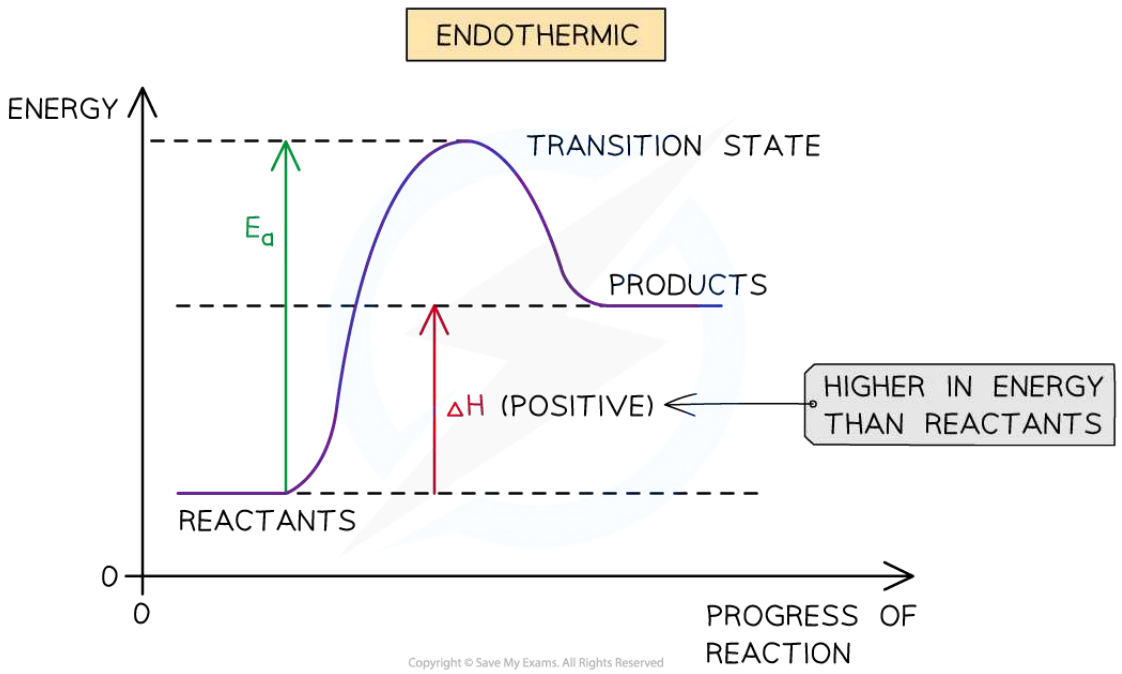
The reaction pathway diagram for exothermic reactions

Endothermic reactions

- If more energy is absorbed to break bonds than is released to form new bonds, this reaction is **endothermic overall**
- The change in energy is positive since the products have more energy than the reactants
- Therefore an endothermic reaction has a **positive** ΔH value, which is shown on the energy level diagrams and in calculations

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The reaction pathway diagram for endothermic reactions.

Exam Tip

You **must** be able to draw these pathway diagrams and label the following parts:

- Reactants
- Products
- Enthalpy change of the reaction, ΔH
- Activation energy. E_a

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Bond Breaking & Bond Forming

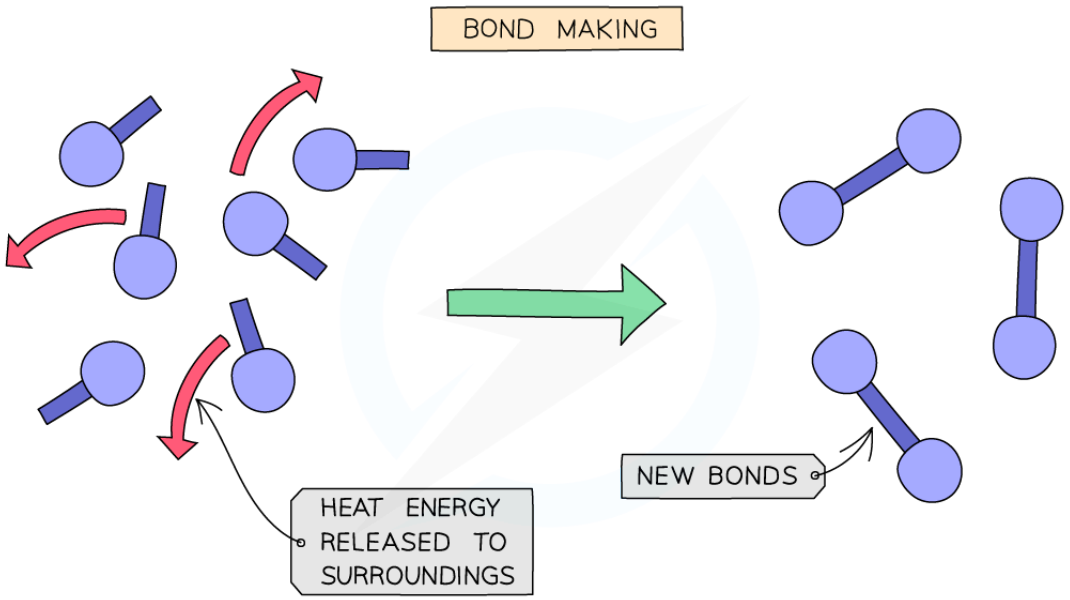
Your notes

Bond Breaking & Bond Forming

- Whether a reaction is endothermic or exothermic depends on the difference between the energy needed to **break** existing bonds and the energy released when the new bonds are **formed**
- **Bond breaking** is always an **endothermic** process as energy needs to be taken in from the surroundings to break the chemical bonds
- **Bond making** is always an **exothermic** process as energy is transferred to the surroundings as the new bond is formed

Exothermic reactions

- If more energy is released than is absorbed, then the reaction is **exothermic**
- More energy is released when new bonds are formed than energy required to break the bonds in the reactants
- The change in energy is negative since the products have less energy than the reactants
- Therefore an exothermic reaction has a **negative** ΔH value



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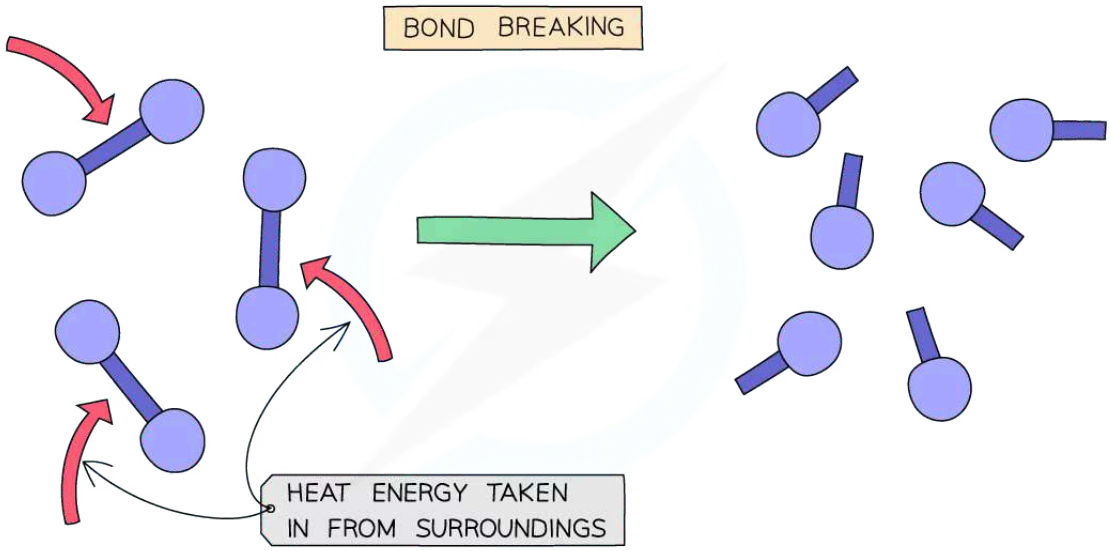
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Making new chemical bonds releases energy which radiates outwards from the reaction to the surroundings in the form of heat

Your notes

Endothermic reactions

- If more energy is absorbed to break bonds than is released to form new bonds, this reaction is **endothermic overall**
- The change in energy is positive since the products have more energy than the reactants
- The symbol ΔH (delta H) is used to show the change in heat energy. H is the symbol for enthaply, which is a measure of the total **heat of reaction** of a chemical reaction
- Therefore an endothermic reaction has a **positive** ΔH value, which is shown on the reaction pathway diagrams and in calculations



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Breaking chemical bonds requires energy which is taken in from the surroundings in the form of heat

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Bond Energy Calculations

Energy of reaction calculations


Your notes

- Each chemical bond has specific bond energy associated with it
- This is the amount of energy required to **break** the bond or the amount of energy given out when the bond is **formed**
- This energy can be used to calculate how much heat would be released or absorbed in a reaction
- To do this it is necessary to know the bonds present in both the reactants and products

Method

- Write a balanced equation if none is present already
- Optional - draw the displayed formula in order to identify the type and number of bonds more easily
- Add together all the bond energies for all the bonds in the reactants – this is the 'energy in'
- Add together the bond energies for all the bonds in the products – this is the 'energy out'
- Calculate the enthalpy change:

Enthalpy change (ΔH) = Energy taken in – Energy given out

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Worked example

Hydrogen and chlorine react to form hydrogen chloride gas:

$$H_2 + Cl_2 \rightarrow 2HCl$$

The table below shows the bond energies.

Bond	Bond energy (kJ/mole)
H–H	436
Cl–Cl	243
H–Cl	432

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Calculate the enthalpy change, ΔH, for the reaction and deduce whether it is exothermic or endothermic.

Answer

Step	Working out
1. Calculate the energy in	Energy in = 436 + 243 = 679 kJ/Mole
2. Calculate the energy out	Energy out = 2 × 432 = 864 kJ/Mole
3. Calculate the enthalpy change	Enthalpy change = 679 − 864 = −185 kJ/Mole
4. Comment on the result and deduce the nature of the reaction	Enthalpy change is a negative number hence energy has been lost to the surrounding and the reaction is exothermic

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Worked example

Hydrogen bromide decomposes to form hydrogen and bromine:

$$2HBr \rightarrow H_2 + Br_2$$

The table below shows the bond energies.

Bond	Bond energy (kJ/mole)
H–Br	366
H–H	436
Br–Br	193

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Calculate the enthalpy change, ΔH, for the reaction and deduce whether it is exothermic or endothermic.

Answer

Step	Working out
1. Calculate the energy in	Energy in = 2 × 366 = 732 kJ/Mole
2. Calculate the energy out	Energy out = 436 + 193 = 629 kJ/Mole
3. Calculate the enthalpy change	Enthalpy change = 732 − 629 = +103 kJ/Mole
4. Comment on the result and deduce the nature of the reaction	Enthalpy change is a positive number hence energy has been absorbed from the surrounding so the reaction is endothermic

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Exam Tip

When answering questions to calculate the enthalpy change using bond energies,, it is helpful to write down a **displayed formula** equation for the reaction **before** identifying the type and number of bonds, to avoid making mistakes. The reaction thus becomes: H-H + Cl-Cl → H-Cl + H-Cl


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Cambridge O Level Chemistry

Your notes

6.1 Physical & Chemical Changes

Contents

Physical & Chemical Changes

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Physical & Chemical Changes


Your notes

Physical & Chemical Changes

Physical change

- Physical changes (such as **melting** or **evaporating**) do not produce any new chemical substances
- These changes are often **easy to reverse**
- Making a mixture from 2 or more substances or dissolving a solute in a solvent are examples of physical changes as no new substances are produced and are usually relatively easy to separate

Chemical change

- During chemical changes (usually referred to as chemical reactions), **new chemical** substances are formed that have very different **properties** to the reactants
- There may be signs that a new substance has formed, such as:
 - A colour change
 - A precipitate being formed
 - Bubbles of gas being produced
- Most chemical reactions are very difficult to **reverse**
- Energy changes also accompany chemical changes and energy can be given out (**exothermic**) or taken in (**endothermic**)
 - The majority of chemical reactions are exothermic with only a small number being endothermic

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Cambridge O Level Chemistry

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6.2 Rate of Reaction

Contents

- Collision Theory
- Explaining Rates Using Collision Theory
- Investigating The Rate of a Reaction
- Interpreting Data

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Collision Theory

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Collision Theory

- Collision theory states that in order for a reaction to occur:
 - The particles must **collide** with each other
 - The collision must have **sufficient energy** to cause a reaction i.e. enough energy to break bonds
- The minimum energy that colliding particles must have to react is known as the **activation energy**
- Collisions which result in a reaction are known as **successful collisions**
 - If they have sufficient energy (i.e. energy greater than the activation energy), they will react, and the collision will be successful
- Not all collisions result in a chemical reaction:
 - Most collisions just result in the colliding particles bouncing off each other
 - Collisions which do not result in a reaction are known as unsuccessful collisions
- Unsuccessful collisions happen when the colliding species do not have enough energy to break the necessary bonds (i.e. they collide with energy less than the activation energy)

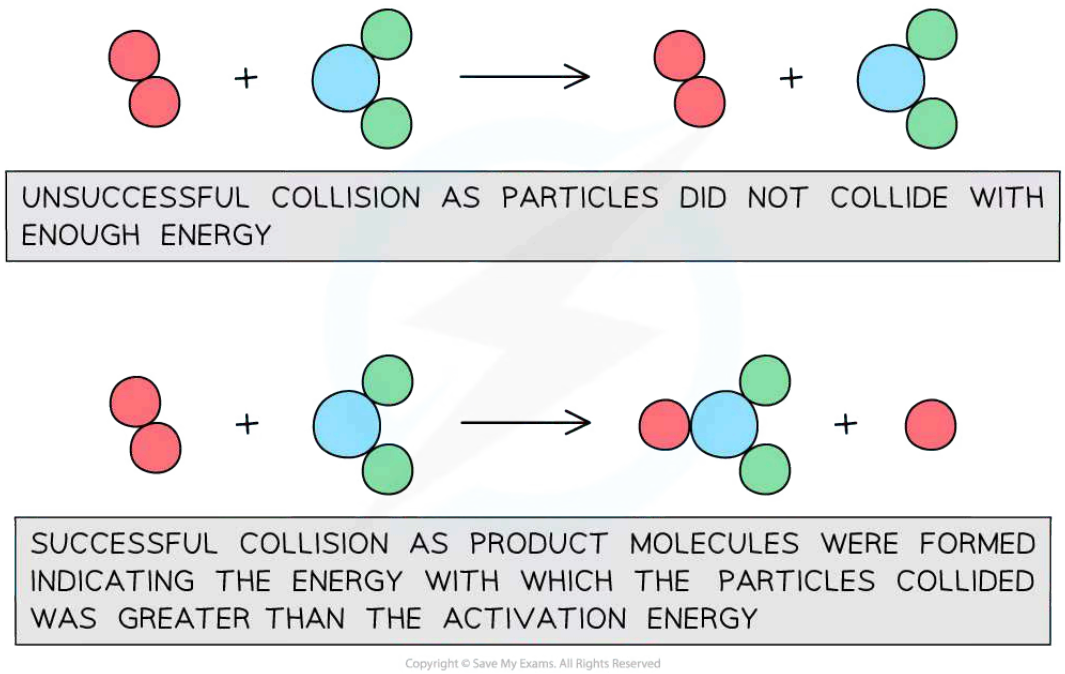


Diagram showing a successful and an unsuccessful collision

- Increasing the number of successful collisions means that a greater proportion of reactant particles collide to form product molecules

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- The number of successful collisions depends on:
 - The **number of particles per unit volume** - more particles in a given volume will produce more frequent successful collisions
 - The **frequency of collisions** - a greater number of collisions per second will give a greater number of successful collisions per second
 - The **kinetic energy of the particles** - greater kinetic energy means a greater proportion of collisions will have an energy that exceeds the activation energy and the more frequent the collisions will be as the particles are moving quicker, therefore, more collisions will be successful
 - The **activation energy** - fewer collisions will have an energy that exceeds higher activation energy and fewer collisions will be successful
- These all have an impact on the rate of reaction which is dependent on the number of successful collisions per unit of time


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Explaining Rates Using Collision Theory

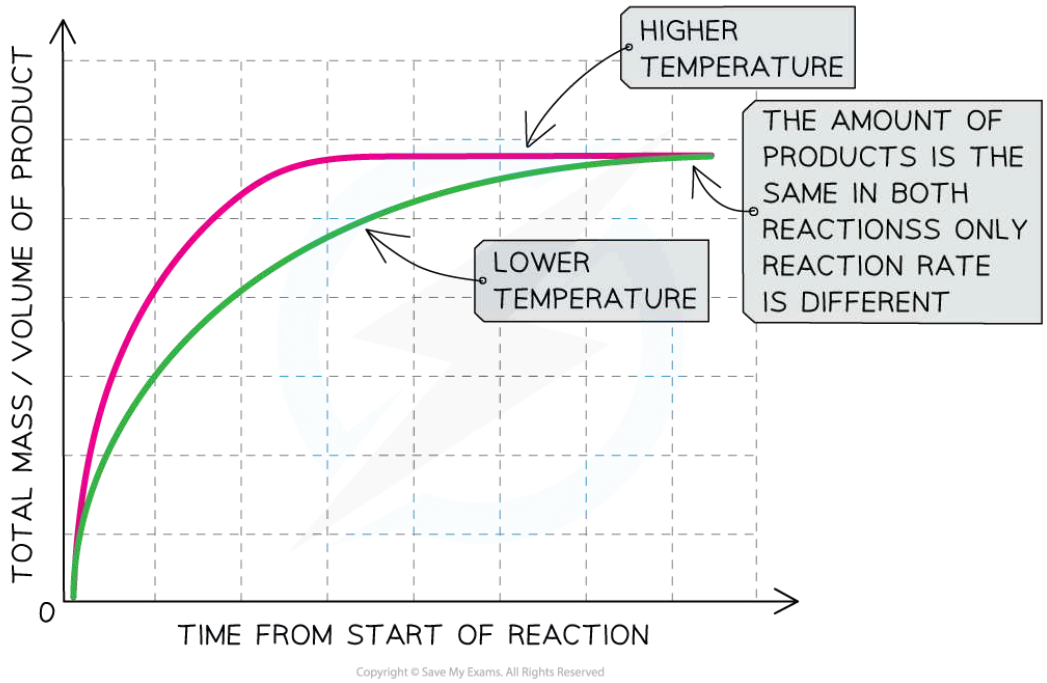
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Explaining Rates Using Collision Theory

Rates of reaction factors

- There are several factors that can affect the rate of a reaction. These are:
 - **Concentration** of the reactants in solution or the **pressure** of reacting gases
 - **Temperature** at which the reaction is carried out
 - **Surface area** of solid reactants
 - The use of a **catalyst**
- Changes in these factors directly influence the rate of a reaction
- It is of **economic interest** to have a higher rate of reaction as this implies a higher rate of production and hence a more efficient and sustainable process

The effect of temperature



- Compared to a reaction at a low temperature, the graph line for the same reaction but at a higher temperature has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased temperature, the rate of reaction will increase**

Explanation:

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- This is because the particles will have more kinetic energy than the required activation energy
- Therefore there will be more frequent collisions **and** a higher proportion of particles have energy greater than the activation energy
- This causes more successful collisions per second, increasing the rate of reaction
- The effect of temperature on collisions is not so straightforward as concentration or surface area; a small increase in temperature causes a large increase in rate
- For aqueous and gaseous systems, a rough rule of thumb is that for every 10 °C increase in temperature, the rate of reaction approximately doubles

Your notes

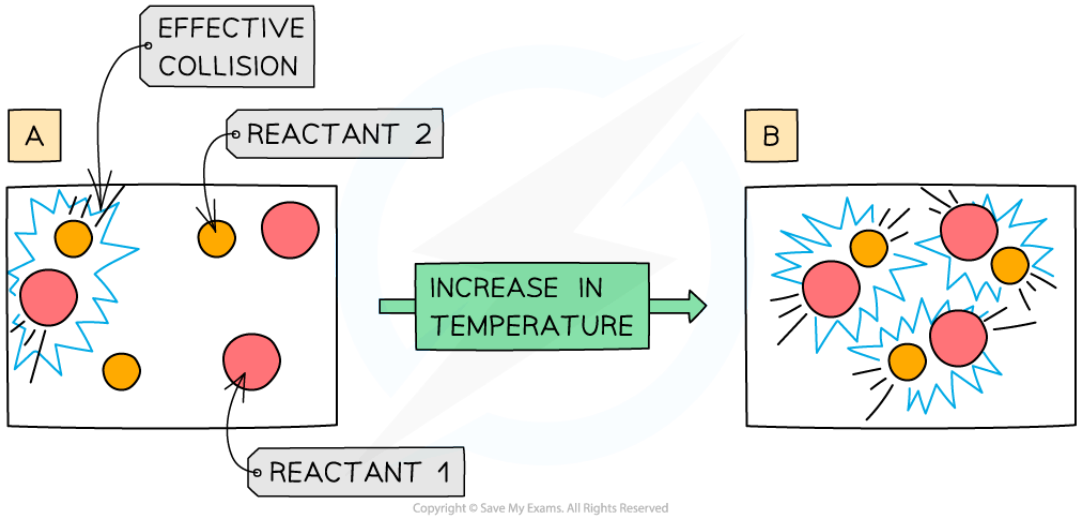
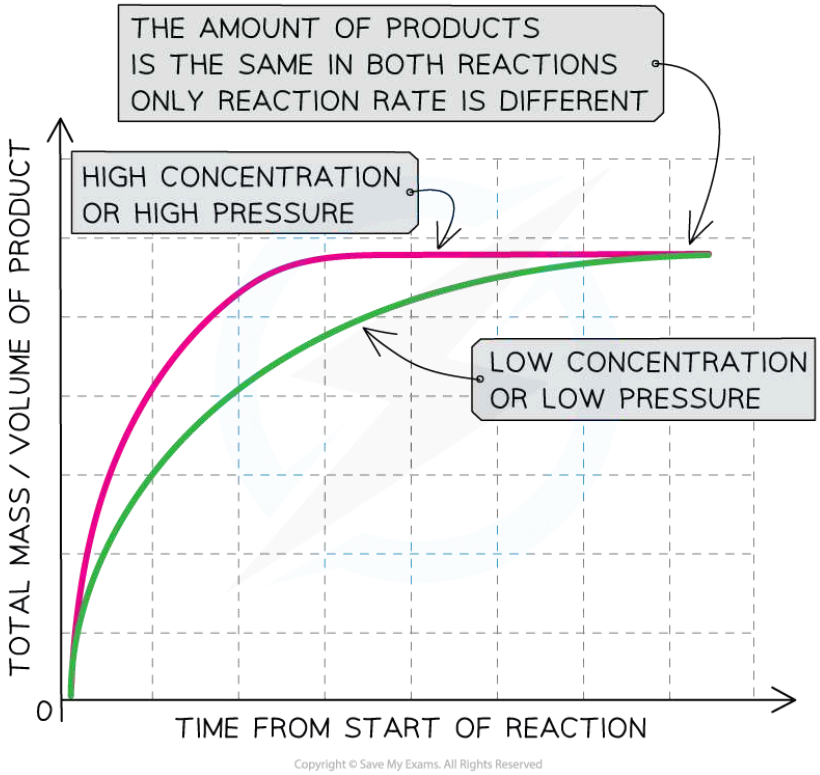


Diagram showing the increased kinetic energy that particles have at higher temperatures

The effect of increased concentration or pressure

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- Compared to a reaction with a reactant at a low concentration, the graph line for the same reaction but at a higher concentration/pressure has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased concentration of a solution or increased pressure of a gas, the rate of reaction will increase**

Explanation:

- **Increasing the concentration of a solution will increase the rate of reaction**
- This is because there will be more reactant particles in a given volume, allowing more frequent and successful collisions per second, increasing the rate of reaction
- For a gaseous reaction, increasing the **pressure** has the same effect as the same number of particles will occupy a smaller space, increasing the concentration
- If you double the number of particles you will double the number of collisions per second
- The number of collisions is **proportional** to the number of particles present

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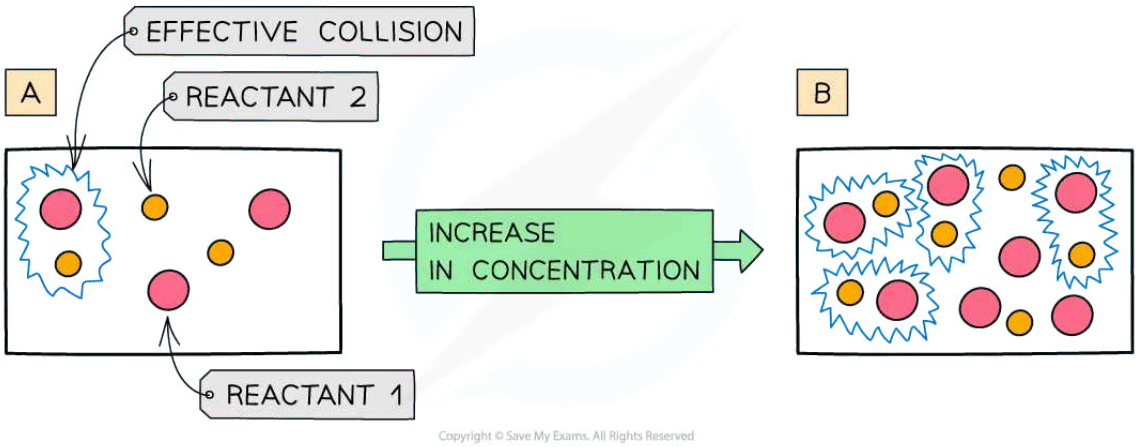
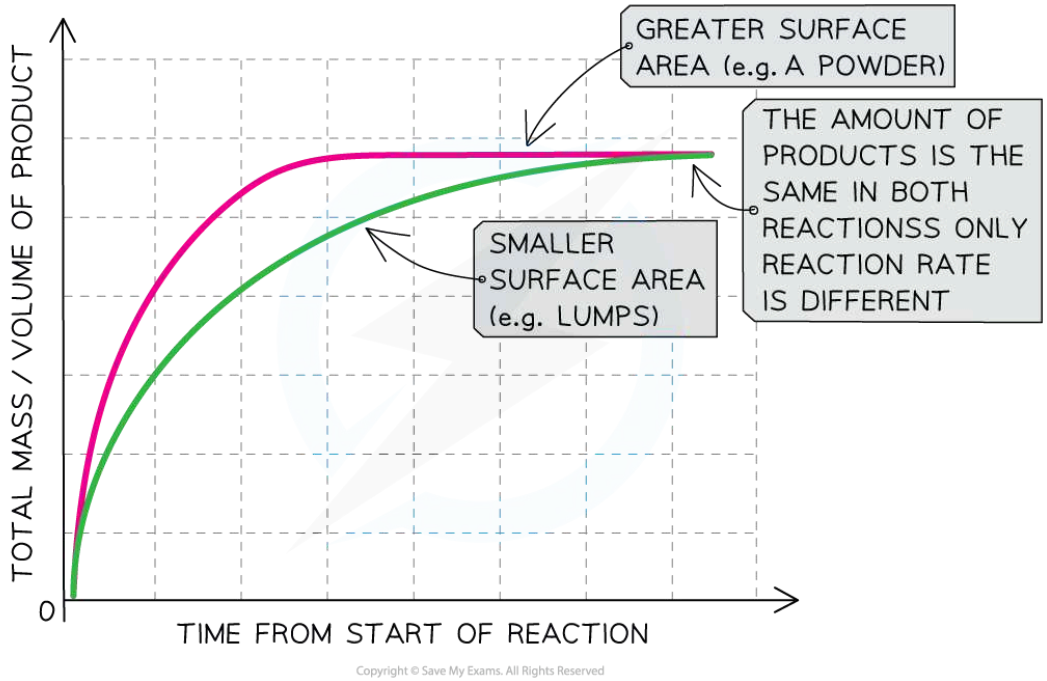


Diagram showing the decrease in space between particles at higher concentrations

Surface Area



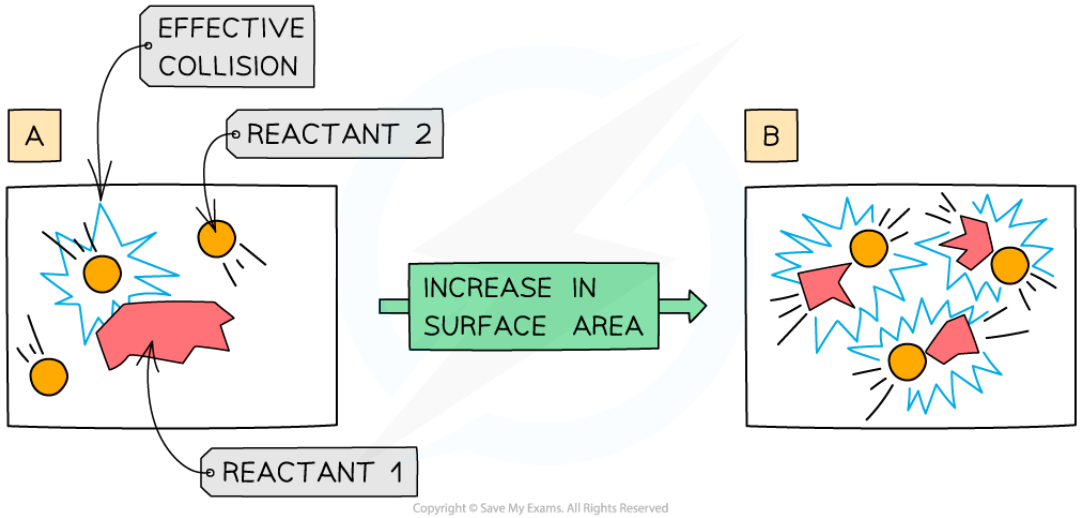
- Compared to a reaction with lumps of reactant, the graph line for the same reaction but with powdered reactant has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased surface area of the solid, the rate of reaction will increase**

Explanation:

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- This is because more **surface area** of the **particles** will be exposed to the other reactant, producing a higher number of collisions per second
- If you double the surface area you will double the number of collisions per second

Your notes

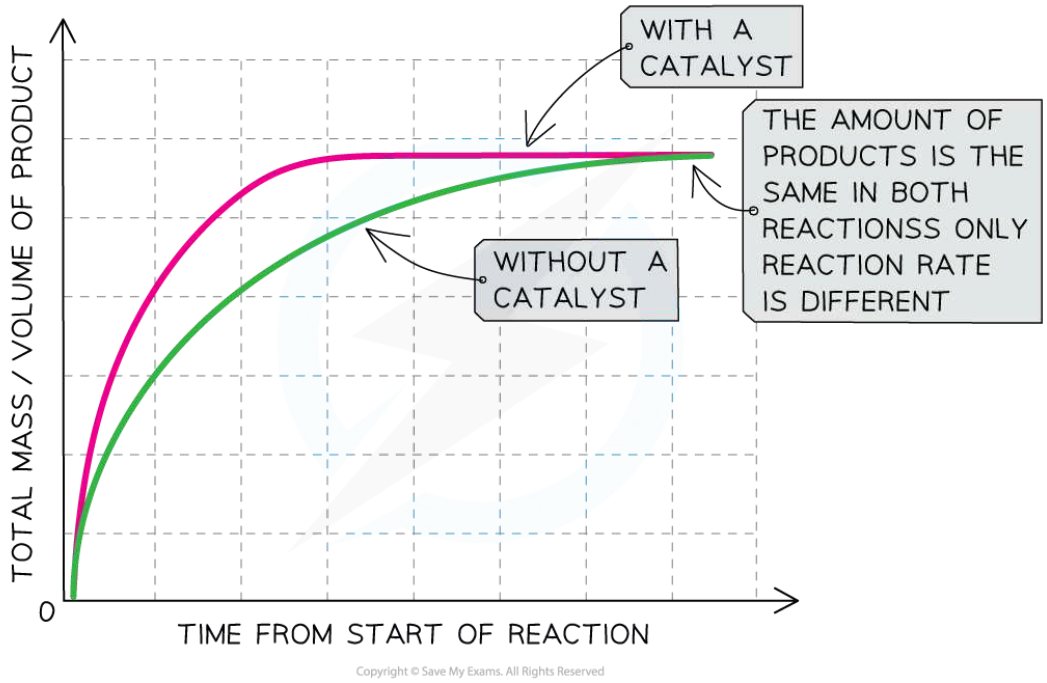


An increase in surface area means more collisions per second

Catalysts

- Catalysts are substances which **speed up** the **rate** of a reaction but are **unchanged** at the end of the reaction
- The mass of a catalyst at the beginning and end of a reaction is the **same** and they do not form part of the equation

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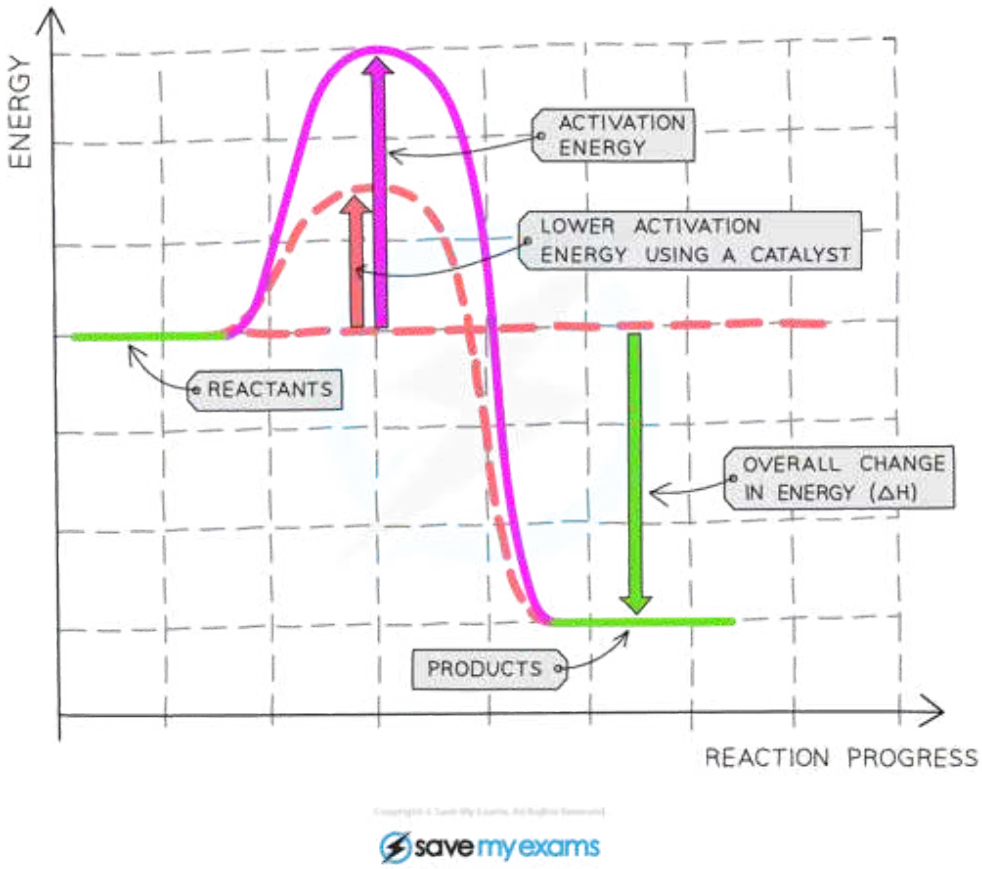
- Compared to a reaction without a catalyst, the graph line for the same reaction but with a catalyst has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with a catalyst, the rate of reaction will increase**

Explanation:

- Different processes require different types of catalysts but they all work on the same principle of providing a **different pathway** for the reaction to occur that has a l**ower activation energy**
- This means a higher proportion of the reactant particles have energy greater than the activation energy and will result in more successful collisions per second
- An important industrial example is iron, which is used to catalyse the **Haber Process** for the production of ammonia
- Iron **beads** are used to increase the **surface area** available for catalysis
- Enzymes are **biological** catalysts, they work best at specific temperature and pH ranges
- Normally only **small amounts** of catalysts are needed to have an effect on a reaction

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Your notes



A catalyst lowers the activation energy of a reaction by providing an alternative reaction pathway

Exam Tip

When answering questions on the effect of concentration on the rate of reaction, you should mention that there are more particles **per unit volume** (usually cm^3) and this causes an increase in the rate of collisions.

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Investigating The Rate of a Reaction


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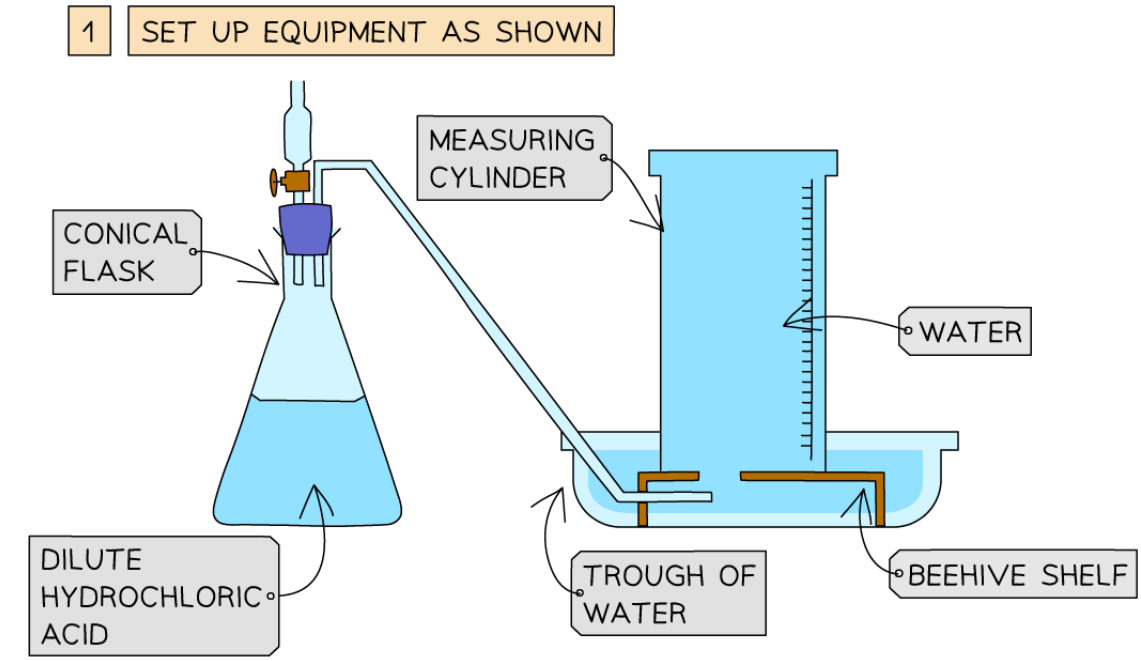
Investigating The Rate of a Reaction

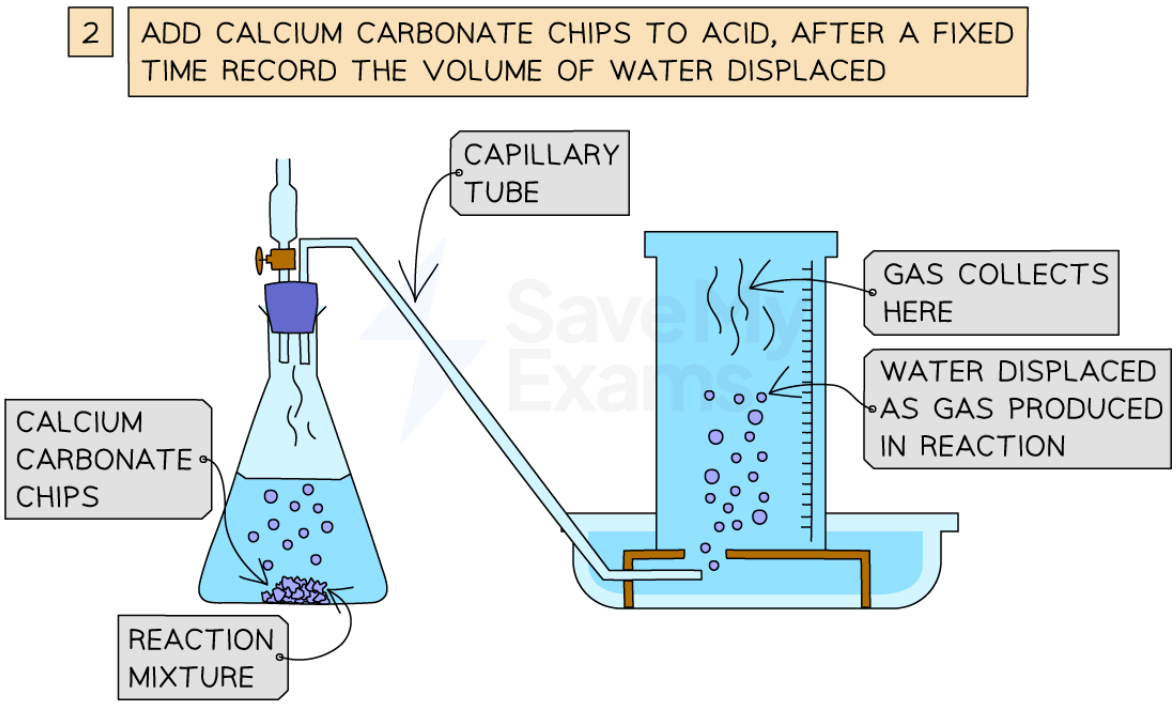
- To measure the **rate of a reaction**, we need to be able to measure either how quickly the reactants are used up or how quickly the products are formed
- The method used for measuring depends on the substances involved
- There are a number of ways to measure a reaction rate in the lab; they all depend on some property that changes during the course of the reaction
- That property is taken to be **proportional** to the concentration of the reactant or product, e.g., colour, mass, volume
 - faster reactions can be easier to measure when the reaction is over, by averaging a collected measurement over the course of the reaction some reaction rates can be measured as the reaction proceeds (this generates more data)
- Three commonly used techniques are:
 - **measuring mass loss on a balance**
 - **measuring the volume of a gas produced**
 - **measuring a reaction where there is a colour change at the end of the reaction**

Investigating the effect of surface area on the rate of reaction

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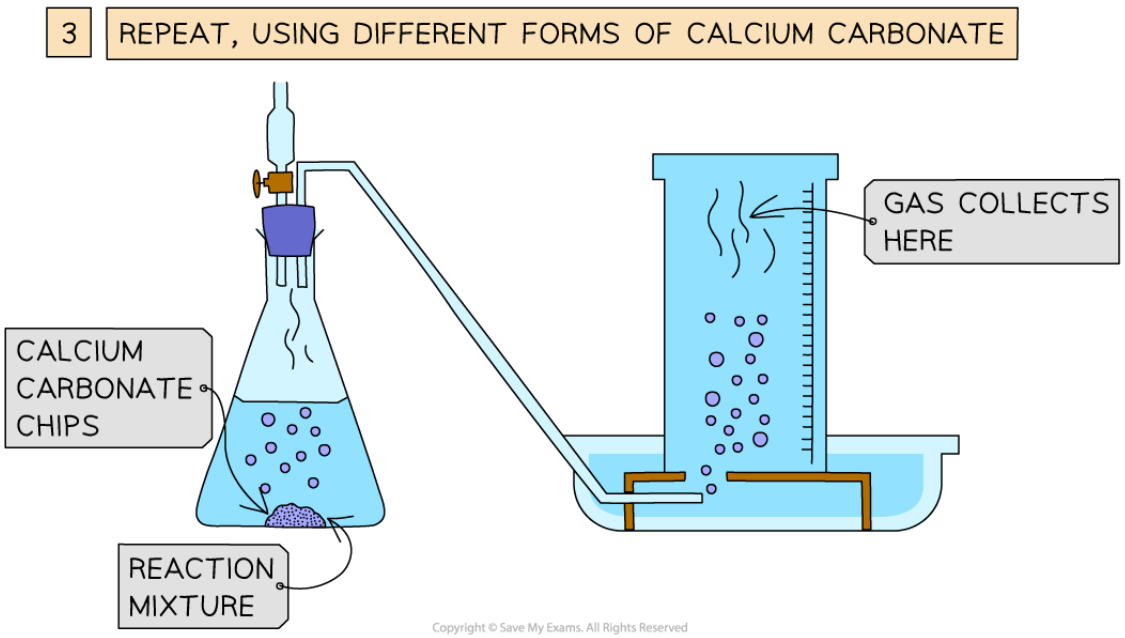


Diagram showing the process of downwards displacement to investigate the effect of the surface area of a solid on the rate of reaction

Method

- Add dilute hydrochloric acid into a conical flask
- Use a delivery tube to connect this flask to a measuring cylinder upside down in a bucket of water (downwards displacement)
- Add calcium carbonate chips into the conical flask and quickly put the bung back into the flask
- Measure the volume of gas produced in a fixed time using the measuring cylinder
- Repeat with different sizes of calcium carbonate chips (lumps, crushed and powdered)

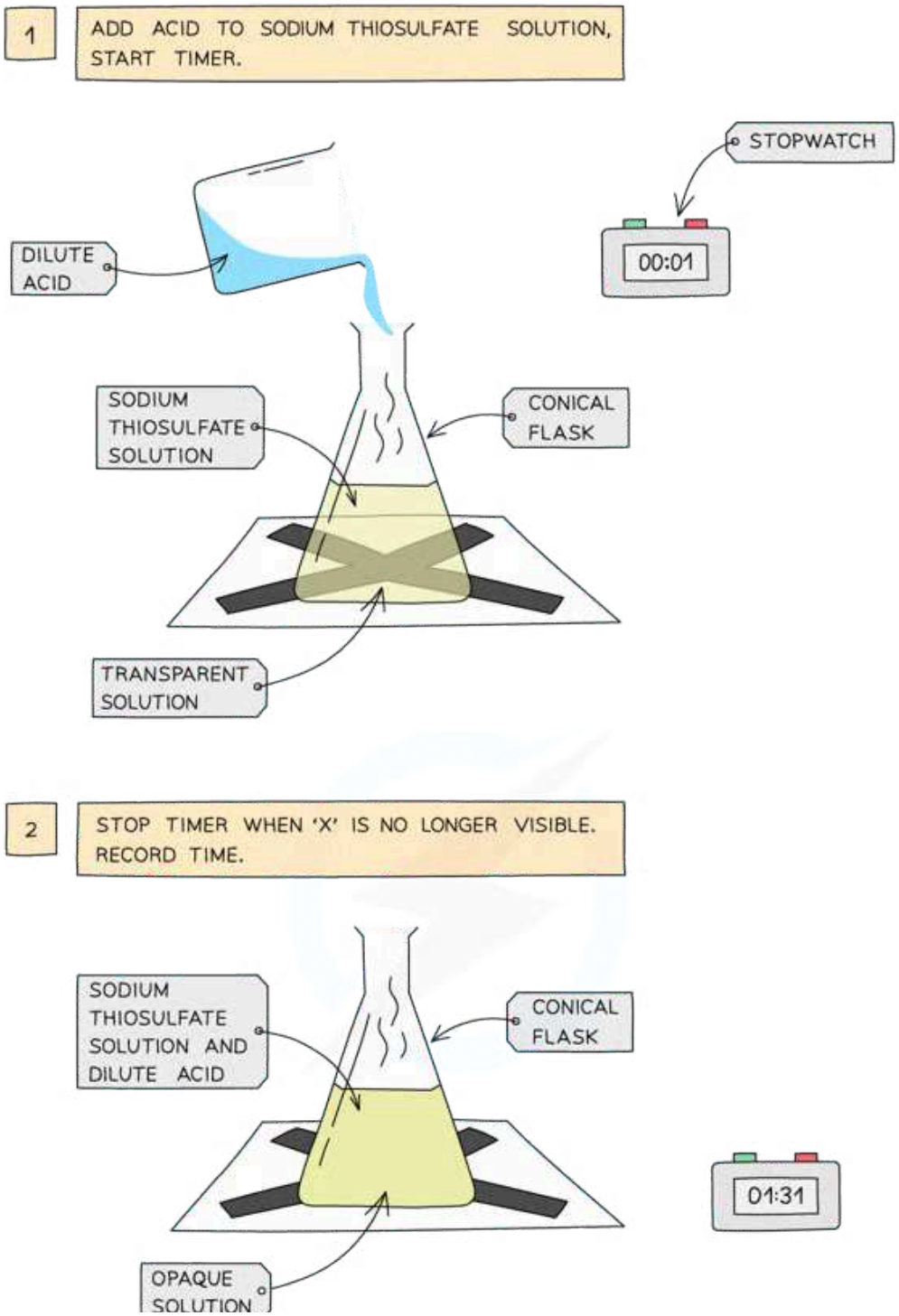
Result

- Smaller sizes of chips cause an increase in the surface area of the solid, so the rate of reaction will increase
- This is because more surface area of the particles will be exposed to the other reactant so there will be more frequent and successful collisions, increasing the rate of reaction

Effect of concentration of a solution on the rate of reaction

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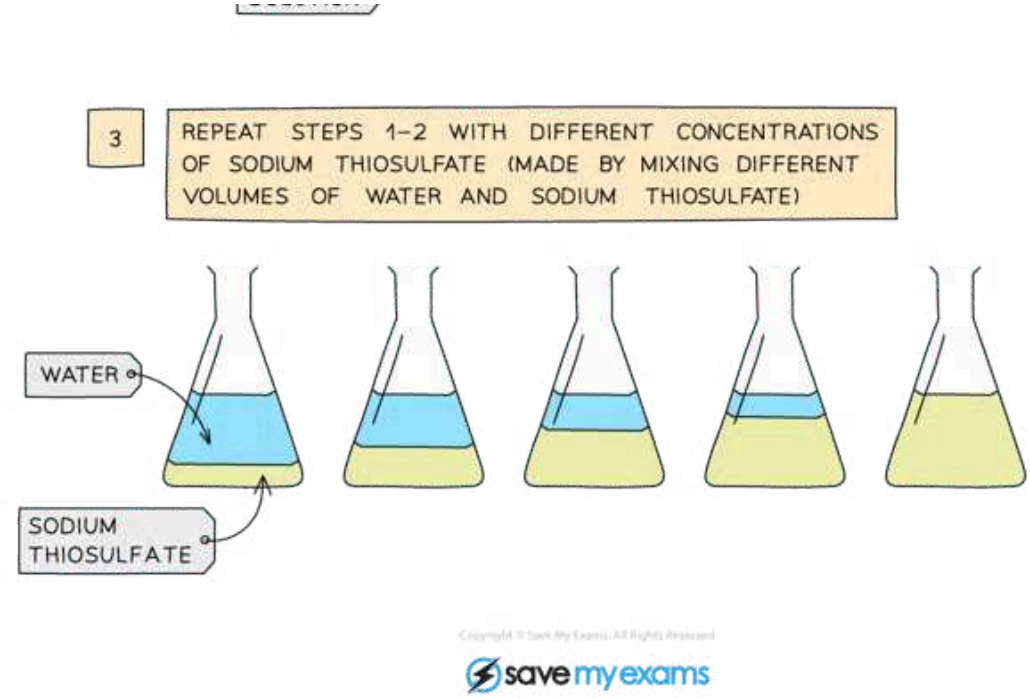


Diagram showing the apparatus needed to investigate the effect of concentration on the rate of reaction

Method

- Measure 50 cm^3 of sodium thiosulfate solution into a flask
- Measure 5 cm^3 of dilute hydrochloric acid into a measuring cylinder
- Draw a cross on a piece of paper and put it underneath the flask
- Add the acid into the flask and immediately start the stopwatch
- Look down at the cross from above and stop the stopwatch when the cross can no longer be seen
- Repeat using different concentrations of sodium thiosulfate solution (mix different volumes of sodium thiosulfate solution with water to dilute it)

Result

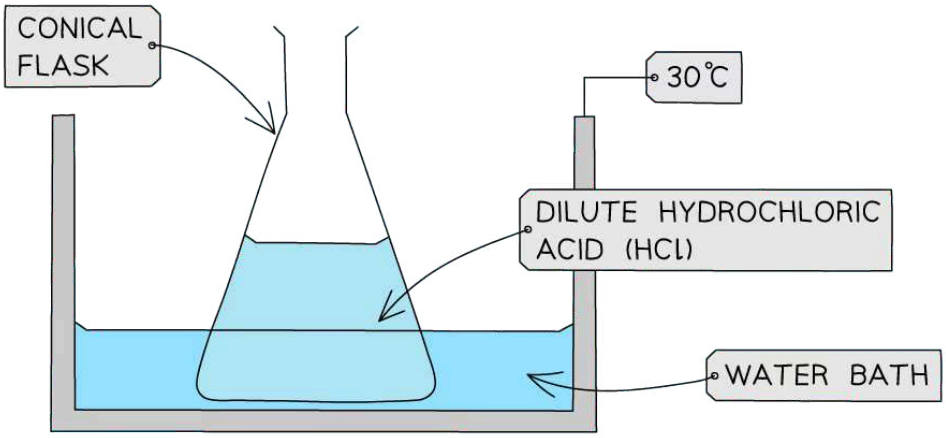
- With an increase in the concentration of a solution, the rate of reaction will increase
- This is because there will be more reactant particles in a given volume, allowing more frequent and successful collisions, increasing the rate of reaction

Effect of temperature on the rate of reaction

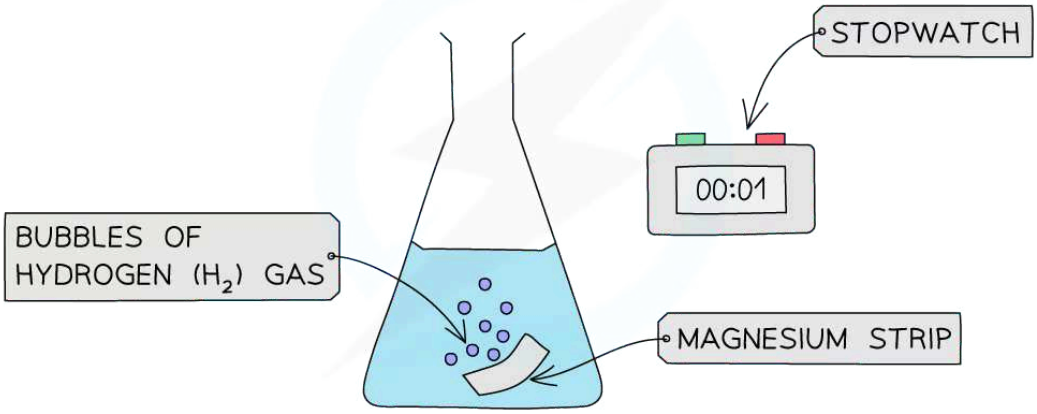
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1 HEAT DILUTE HYDROCHLORIC ACID TO A DESIRED TEMPERATURE



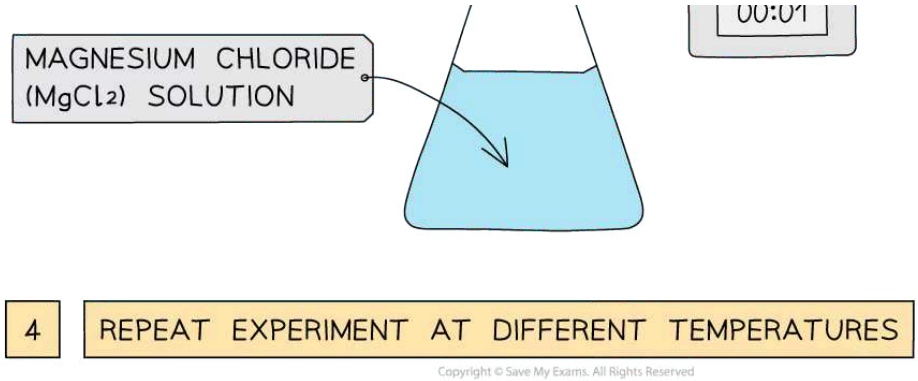
2 ADD A STRIP OF MAGNESIUM TO THE DILUTE HCl AND START A STOPWATCH



3 WHEN THE MAGNESIUM STRIP HAS DISSOLVED, STOP THE STOPWATCH AND RECORD THE TIME



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Diagram showing the apparatus needed to investigate the effect of temperature on the rate of reaction

Method

- Dilute hydrochloric acid is heated to a set temperature using a water bath
- Add the dilute hydrochloric acid into a conical flask
- Add a strip of magnesium and start the stopwatch
- Stop the time when the magnesium fully reacts and disappears
- Repeat at different temperatures and compare results

Result

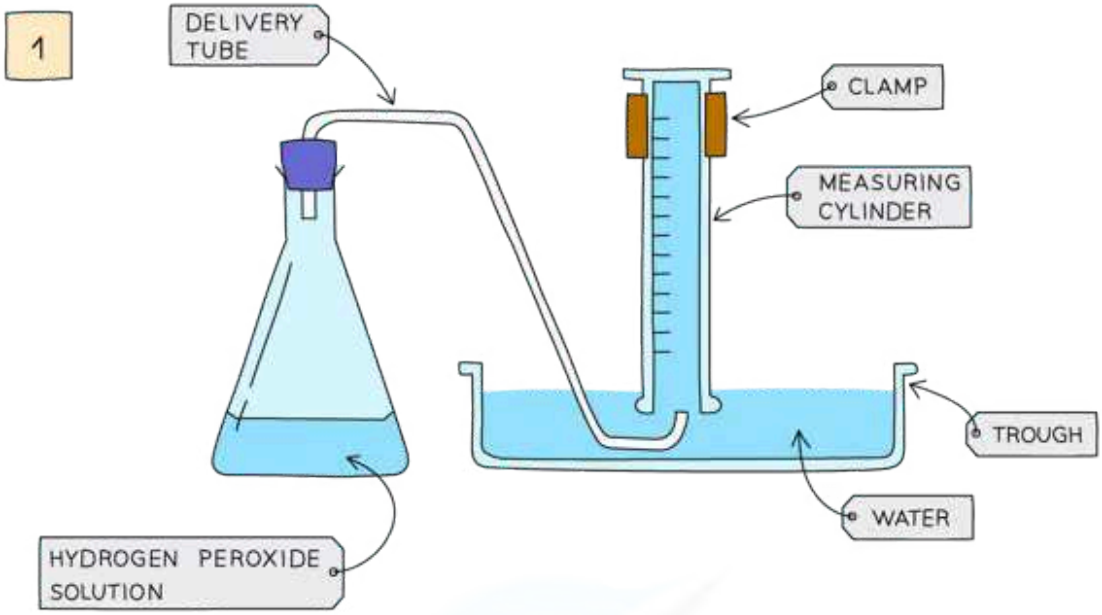
- With an increase in the temperature, the rate of reaction will increase
- This is because the particles will have more kinetic energy than the required activation energy, therefore more frequent and successful collisions will occur, increasing the rate of reaction

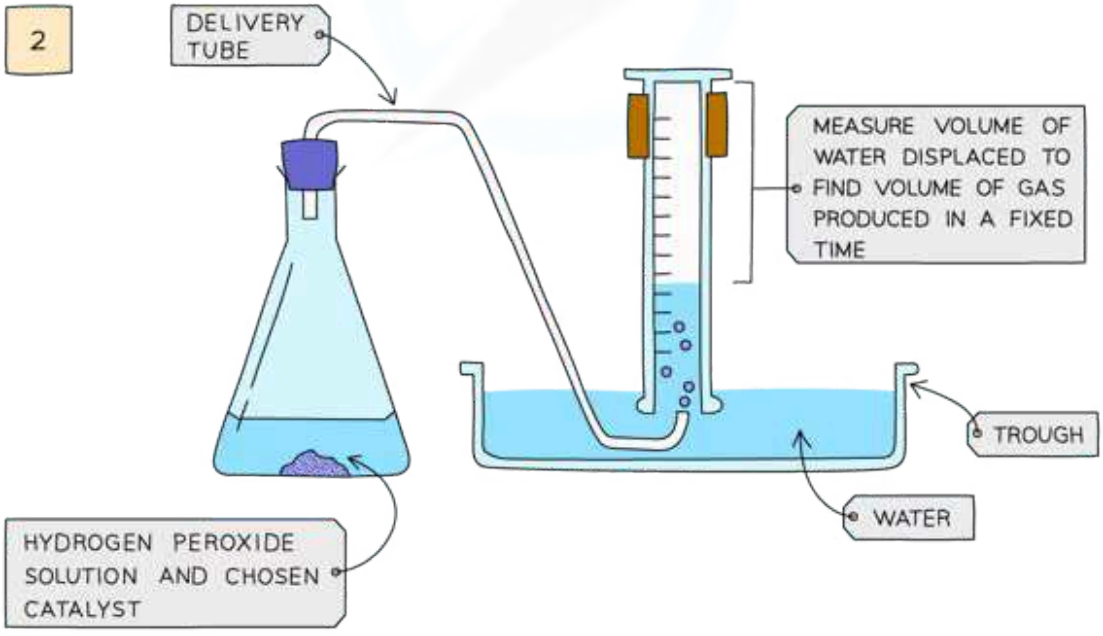
Effect of a catalyst on the rate of reaction

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Diagram showing the apparatus needed to investigate the effect of a catalyst on the rate of reaction

Method

- Add hydrogen peroxide into a conical flask
- Use a delivery tube to connect this flask to a measuring cylinder upside down in a tub of water (downwards displacement)
- Add the catalyst manganese(IV) oxide into the conical flask and quickly place the bung into the flask
- Measure the volume of gas produced in a fixed time using the measuring cylinder
- Repeat experiment without the catalyst of manganese(IV) oxide and compare results

Result

- Using a catalyst will increase the rate of reaction
- The catalyst will provide an alternative pathway requiring lower activation energy so more colliding particles will have the necessary activation energy to react
- This will allow more frequent and successful collisions, increasing the rate of reaction

Monitoring changes in mass

- Many reactions involve the production of a gas which will be released during the reaction
- The gas can be collected and the volume of gas monitored as per some methods above
- Alternatively, the reaction can be performed in an **open flask** on a balance to measure the loss in mass of reactant
- Cotton wool is usually placed in the mouth of the flask which allows gas out but prevents any materials from being ejected from the flask (if the reaction is vigorous)

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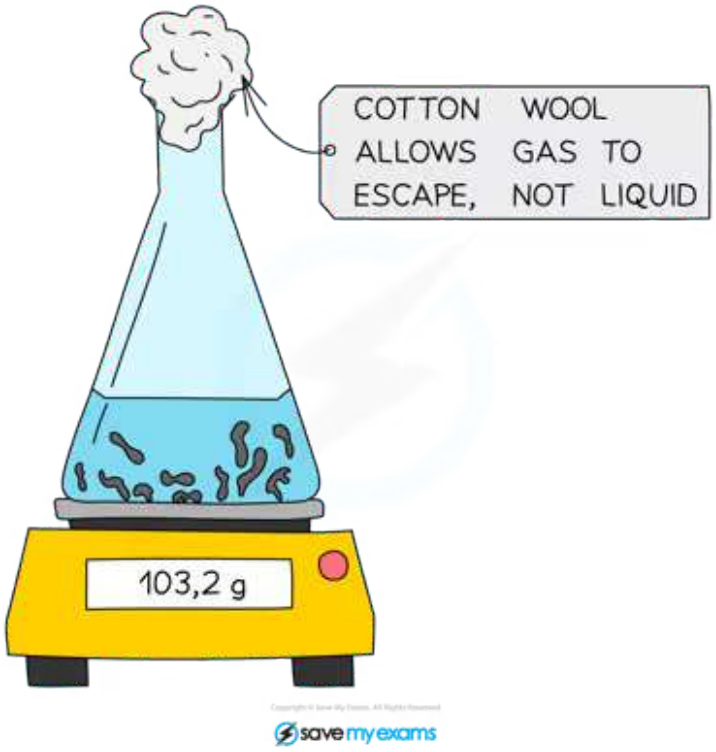


Diagram showing the set-up for measuring the rate of reaction by loss in mass

- This method is not suitable for hydrogen and other gases with a small relative formula mass, M_r as the loss in mass may be too small to measure

Exam Tip

There are many different methods of investigating the rate of reaction. Another method of gas collection you may see uses a gas syringe.

You may be required to devise and evaluate methods of investigating rates of reaction.

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Evaluating Investigations of Rates of Reactions

Your notes

- When investigating rates of reaction, there are a number of different methods that can be used to carry out the same investigation
- Evaluating what is the best method to use is part of good experimental planning and design
- This means appreciating some of the advantages and disadvantages of the methods available

Table showing some Examples of Advantages and Disadvantages of Methods of Investigating Rates of Reaction

Method Choice	Advantage	Disadvantage
Disappearing cross experiment	◦ Simple experiment with no specialist equipment required	◦ Difficult to determine when cross is obscured ◦ Different people will determine the cross to have disappeared at different levels of cloudiness
Collecting gas using a gas syringe	◦ All the gas collected is from the reaction ◦ Gas syringes are easy to set up and keep gas dry	◦ Gas syringes are expensive and fragile ◦ The pistons can stick ◦ Limited volume can be collected ◦ Gas can be lost if the bung is not put onto the top of the flask quickly
Collecting gas using an inverted measuring cylinder	◦ The experiment uses common lab equipment	◦ The delivery tube beneath the measuring cylinder can come away ◦ It can be difficult to read the scale especially when the reaction is vigorous with a large number of bubbles produced rapidly ◦ Gas can be lost if the bung is not put onto the top of the flask quickly
Measuring the loss of mass of reactants	◦ Simple experiment to carry out with no specialist equipment	◦ Not suitable for gases with a small Mr

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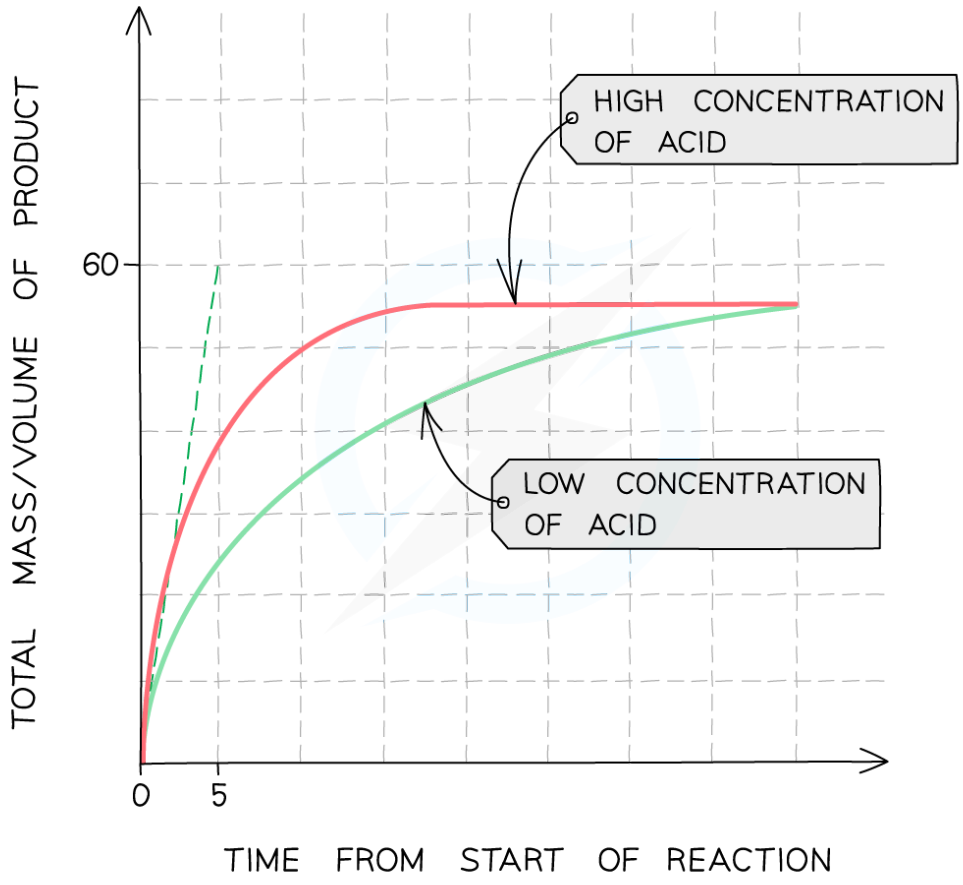
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Interpreting Data

Your notes

Interpreting Data

- Data recorded in rate studies is used to **plot graphs** to calculate the rate of a reaction
- Plotting a graph until the completion of the reaction shows how the rate changes with time
- Over time the rate of reaction slows as the reactants are being used up so the line becomes less steep and eventually becomes **horizontal**, indicating the reaction has finished
- You can plot more than one run of a variable on the same graph making it easier to see how the variable influences the rate
 - For example, plotting the effect of concentration on a reaction between the acid and marble chips



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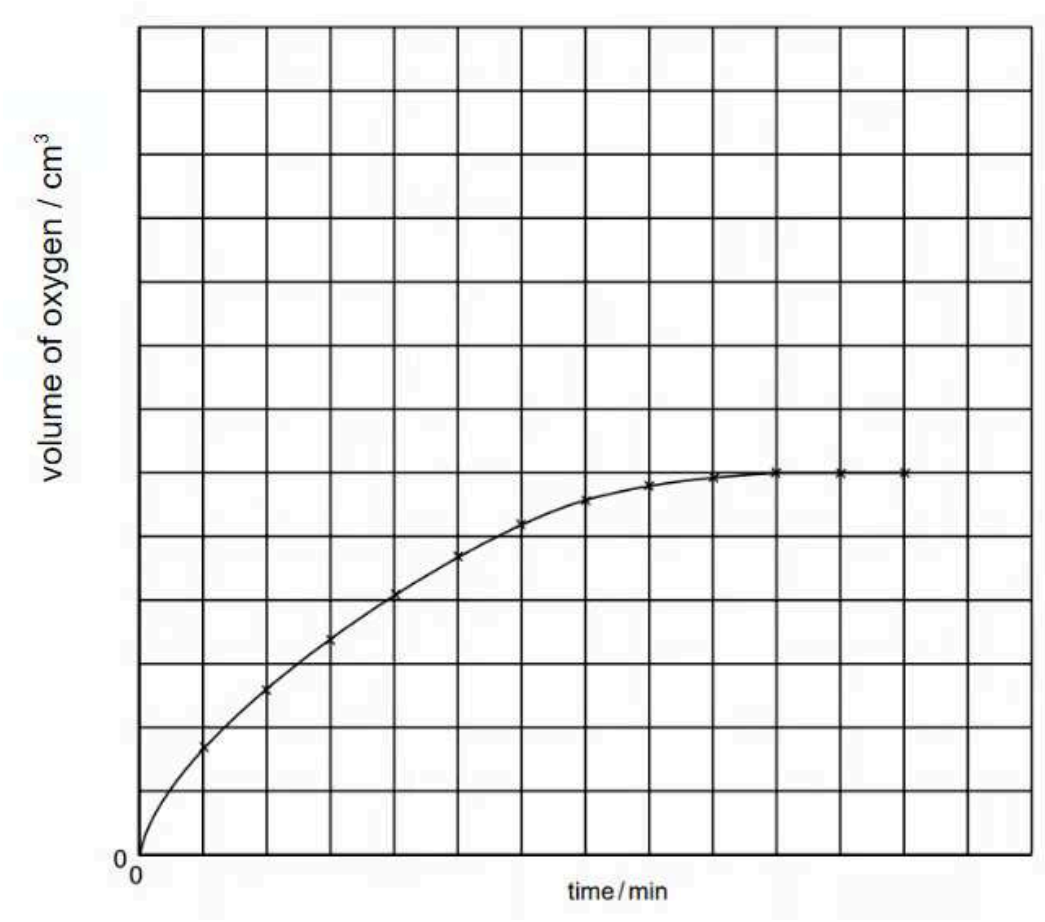
- The steeper the curve, the faster the rate of the reaction
- The curve is steepest initially so the rate is quickest at the beginning of the reaction
- As the reaction progresses, the concentration of the reactants decreases and the rate decreases shown by the curve becoming less steep
- When one of the reactants is used up, the reaction stops, the rate becomes zero and the curve levels off to a horizontal line
- The amount of product formed in a reaction is determined by the limiting reactant:
 - If the amount of limiting reactant increases, the amount of product formed increases
 - If the amount of the reactant in excess increases, the amount of product remains the same
- Drawing a **tangent** to the slope allows you to show the gradient at any point on the curve
- The steeper the slope, the quicker the rate of reaction
- The volume of a gaseous product would increase to a maximum over time, so the line levels out indicating the reaction is over
- Since the volume and mass would be proportional, this could also be a graph of the mass of product versus time


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Worked example

0.2 g of manganese(IV) oxide was added to 25 cm^3 of 0.1 mol/dm^3 hydrogen peroxide solution. The volume of oxygen produced every minute was recorded and the results are shown on the graph.



The experiment was repeated using the same mass of manganese(IV) oxide and at the same temperature but using 25 cm^3 of 0.2 mol/dm^3 of hydrogen peroxide solution.

Sketch the curve for the results of this experiment on the same grid.

Answer

Step 1 - Deduce how the initial gradient will be different from the original graph

The hydrogen peroxide solution is twice as concentrated so the rate of reaction will be greater and the initial gradient will be steeper

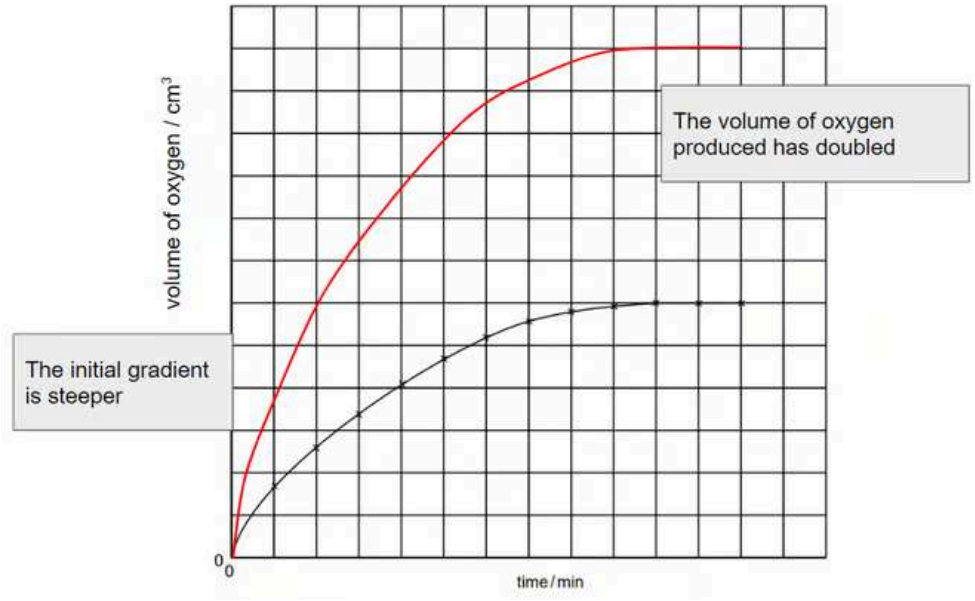
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Step 2 - Deduce how much product will be formed compared to the original experiment

The amount of hydrogen peroxide determines the amount of oxygen produced. In the 2nd experiment, there are twice as many hydrogen peroxide molecules in the same volume so the amount of oxygen gas produced will be doubled

Step 3 - Sketch the graph



Calculating the Rate of Reaction at a Particular Point

- To do this you need to find the **gradient of the curve at that point**
- To do this a **tangent is drawn to the curve** and then the gradient of the tangent calculated

Your notes

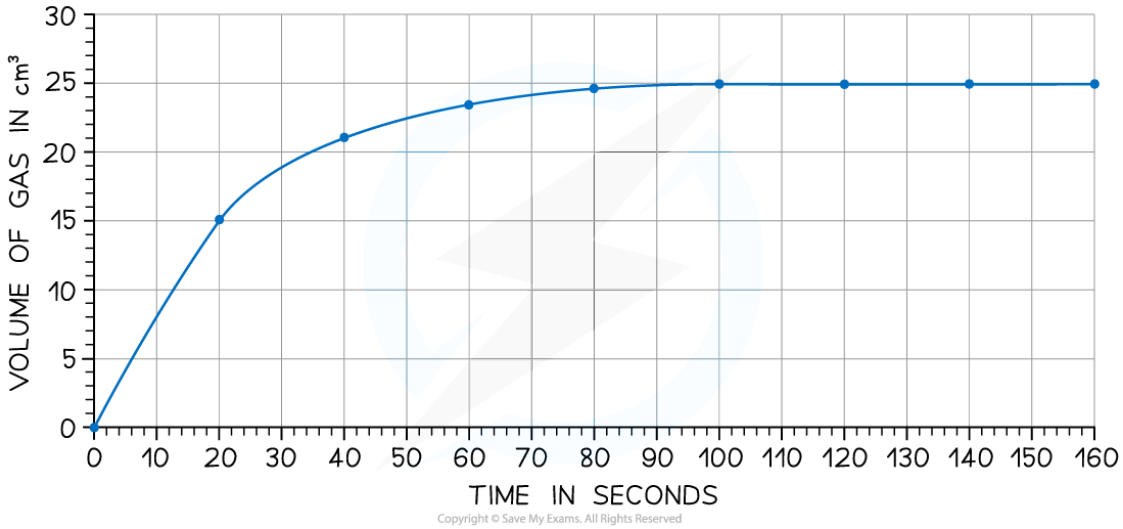
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Worked example

Iodine and methanoic acid react in aqueous solution.

$$I_2\,(aq) + HCOOH\,(aq) \rightarrow 2I^-\,(aq) + 2H^+\,(aq) + CO_2\,(g)$$

The rate of reaction can be found by measuring the volume of carbon dioxide produced per unit time and plotting a graph as shown:



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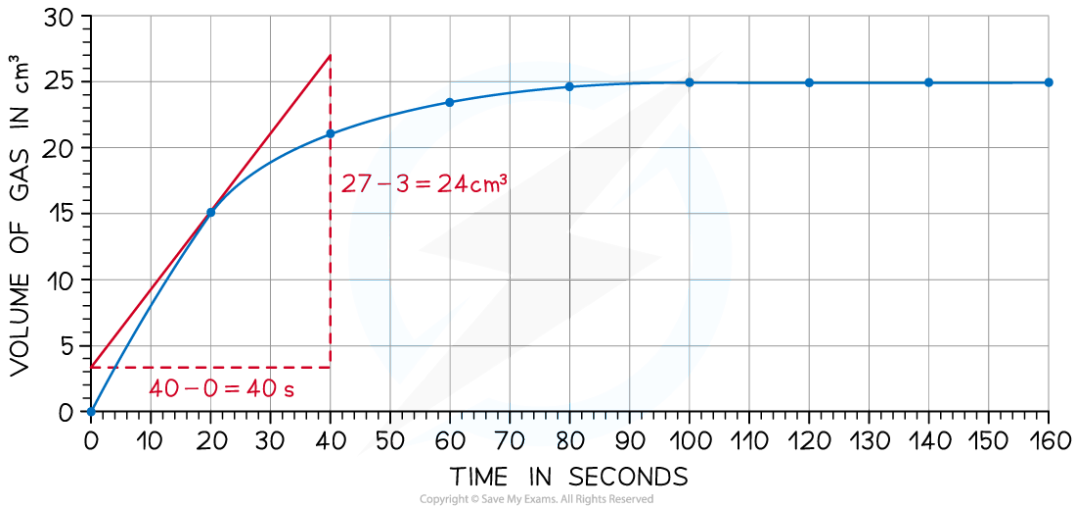
Calculate the rate of reaction at 20 seconds.

Answer

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- Draw a tangent to the curve at 20 seconds:



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- Complete the triangle and read off the values of x and y
- Determine the gradient of the line using change in y / change in x
- Rate of reaction = 24 ÷ 40 = 0.60 cm^3/s

Exam Tip

If the amount of reactant used up is being monitored, then the graph will fall with the steepest gradient at the start, becoming less steep until it levels off to a horizontal line.

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Cambridge O Level Chemistry

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6.3 Reversible Reactions & Equilibrium

Contents

- Reversible Reactions
- Equilibrium
- The Haber Process
- The Contact Process

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Reversible Reactions

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Reversible Reactions

Reversible reactions

- Some reactions go to **completion**, where the reactants are used up to form the product molecules and the reaction stops when all of the reactants are used up
- In reversible reactions, the product molecules can themselves react with each other or decompose and form the reactant molecules again
- It is said that the reaction can occur in **both directions**: the forward reaction (which forms the products) and the reverse reaction(which forms the reactants)

Chemical equations for reversible reactions

- When writing chemical equations for reversible reactions, two arrows are used to indicate the forward and reverse reactions
- Each one is drawn with just half an arrowhead – the top one points to the right, and the bottom one points to the left: ⇌

Example

- The reaction for the **Haber process** which is the production of ammonia from hydrogen and nitrogen:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

Hydrated and anhydrous salts

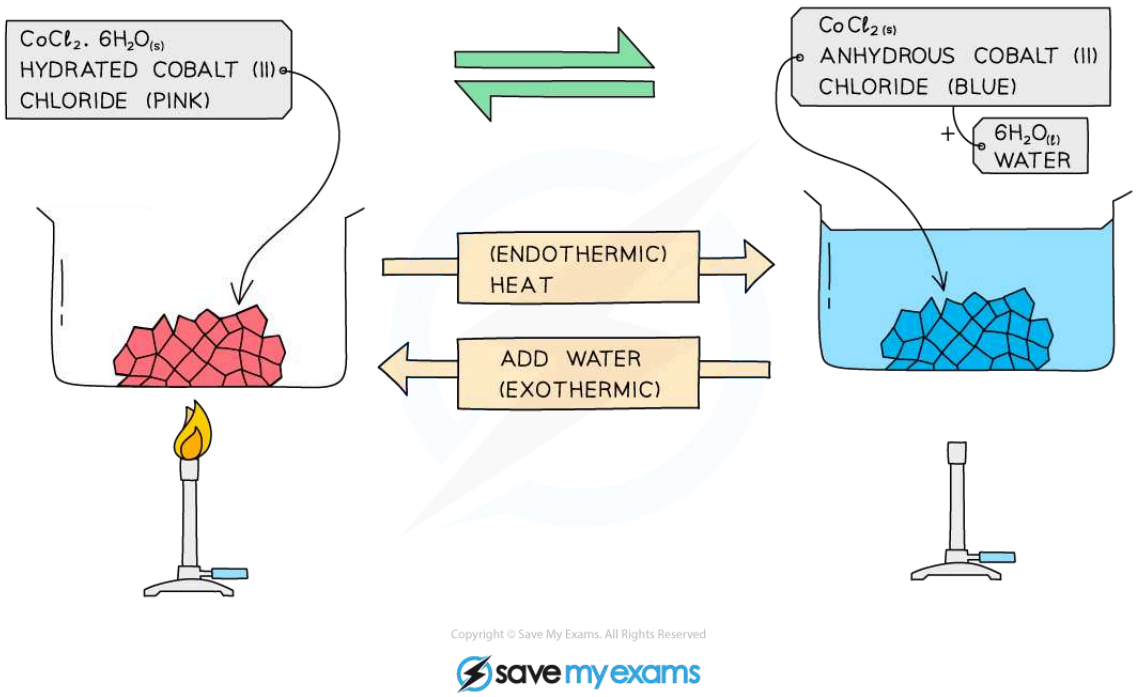
- Hydrated salts are salts that contain **water of crystallisation** which affects their molecular shape and colour
- Water of crystallisation is the water that is stoichiometrically included in the structure of some salts during the crystallisation process
- A common example is copper(II) sulfate which crystallises forming the salt copper(II) sulfate pentahydrate, $CuSO_4.5H_2O$
- Water of crystallisation is indicated with a dot written in between the salt molecule and the surrounding water molecules
- Anhydrous salts are those that have lost their water of crystallisation, usually by heating, in which the salt becomes **dehydrated**

Dehydration of hydrated cobalt(II) chloride:

hydrated cobalt(II) chloride ⇌ anhydrous cobalt(II) chloride + water

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Diagram showing the dehydration of hydrated cobalt(II) chloride

Hydration of cobalt(II) chloride

- When anhydrous **blue** cobalt(II) chloride crystals are added to water they turn **pink** and the reaction is **reversible**
- When the cobalt(II) chloride crystals are heated in a test tube, the **pink** crystals turn back to the **blue** colour again as the water of crystallisation is lost
- The form of cobalt(II) chloride in the crystals that are pink is known as hydrated cobalt(II) chloride because it contains water of crystallisation
- When hydrated cobalt(II) chloride is heated, it loses its water of crystallisation and turns into anhydrous cobalt(II) chloride:

$$CoCl_2.6H_2O\,(s) \rightleftharpoons CoCl_2\,(s) + 6H_2O\,(l)$$

Exam Tip

The hydration of $CoCl_2$ and $CuSO_4$ are chemical tests which are commonly used to detect the presence of water. You should remember the equations and colour changes:

- $CoCl_2 + 6H_2O \rightleftharpoons CoCl_2.6H_2O$ **Blue to pink**
- $CuSO_4 + 5H_2O \rightleftharpoons CuSO_4.5H_2O$ **White to blue**

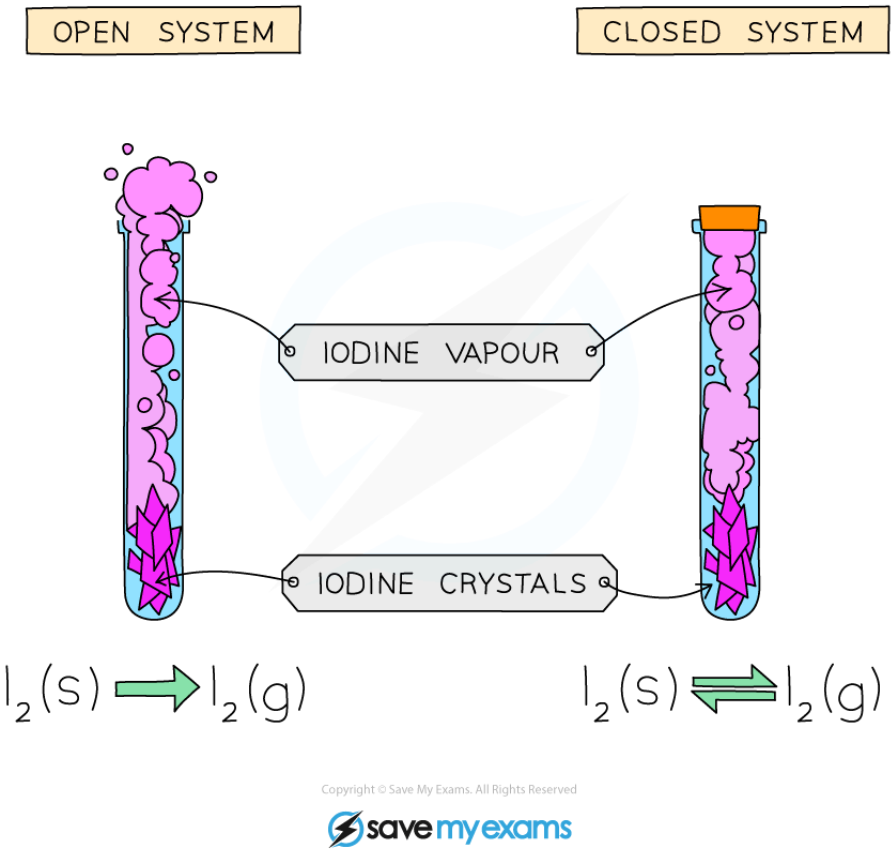
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Equilibrium

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The Concept of Equilibrium

- We have already seen that a reversible reaction is one that occurs in **both** directions
- When during the course of reaction, the **rate of the forward reaction equals the rate of the reverse reaction**, then the overall reaction is said to be in a state of **equilibrium**
- Equilibrium is **dynamic** e.g. the molecules on the left and right of the equation are **changing** into each other by chemical reactions constantly and at the same rate
- The concentration of reactants and products remains **constant** (given there is no other change to the system such as temperature and pressure)
- It only occurs in a **closed system** so that none of the participating chemical species are able to leave the reaction vessel



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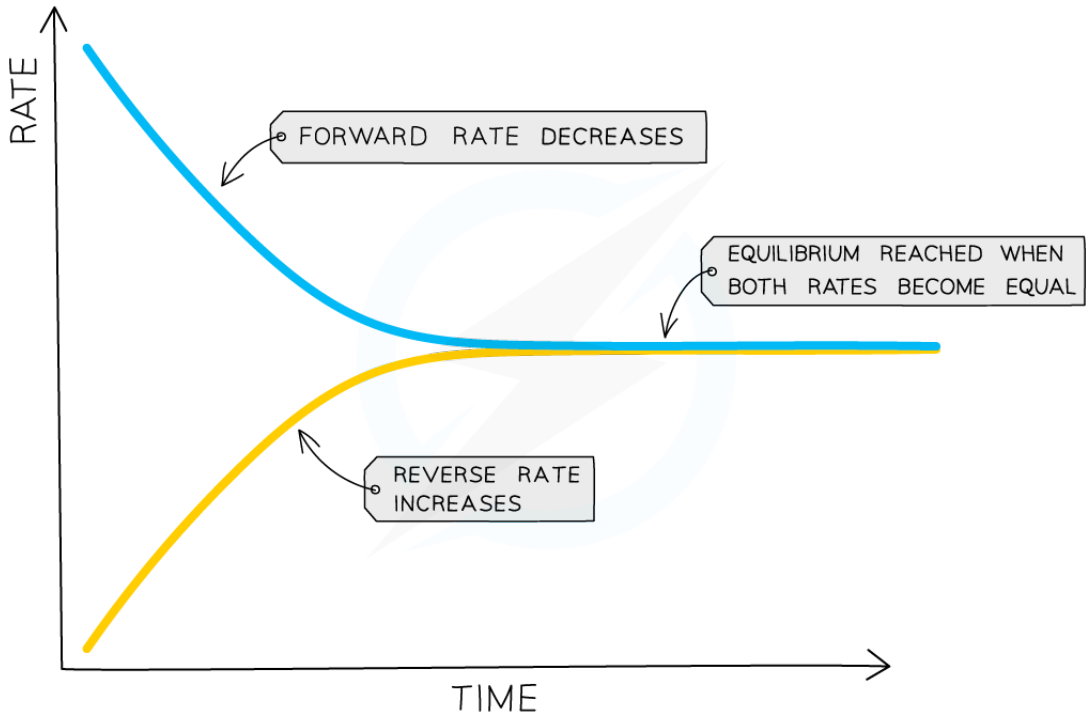
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Equilibrium can only be reached in a closed vessel which prevents reactants or products from escaping the system

Your notes

- An example of dynamic equilibrium is the reaction between H_2 and N_2 in the Haber process
- When only nitrogen and hydrogen are present at the beginning of the reaction, the rate of the forward reaction is at its **highest,** since the **concentrations** of hydrogen and nitrogen are at their **highest**
- As the reaction proceeds, the concentrations of hydrogen and nitrogen gradually decrease, so the rate of the forward reaction will **decrease**
- However, the concentration of ammonia is gradually increasing and so the rate of the **backward** reaction will increase (ammonia will decompose to reform hydrogen and nitrogen)
- Since the two reactions are interlinked and none of the gas can escape, the rate of the forward reaction and the rate of the backward reaction will eventually become **equal** and equilibrium is reached:

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$



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Diagram showing when the rates of forward and backward reactions become equal

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- Equilibrium position refers to the relationship between the concentration of reactants and products at the equilibrium state
- When the position of equilibrium shifts to the **left**, it means the concentration of **reactant** increases
- When the position of equilibrium shifts to **right**, this means the concentration of **product** increases

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Le Chatelier's Principle

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- The relative amounts of all the reactants and products at equilibrium depend on the conditions of the reaction
- This balance is framed in an important concept known as Le Chatelier's Principle, named after Henri Le Chatelier who was a French military engineer in the 19th century
- This principle states that when a change is made to the conditions of a system at equilibrium, the system automatically moves to **oppose** the change
- The principle is used to predict changes to the position of equilibrium when there are changes in **temperature**, **pressure** or **concentration**
- Knowing the energy changes, states and concentrations involved allows us to use the principle to manipulate the outcome of reversible reactions
- For example, if the pressure is increased, the position of equilibrium moves in the direction which has the smallest amount of gaseous molecules
- The position of equilibrium is said to shift to the **right** when the **forward** reaction is favoured and there is an increase in the amount of **products** formed
- The position of equilibrium is said to shift to the **left** when the **reverse** reaction is favoured and there is an increase in the amount of **reactants** formed

The Effect of Temperature on Equilibrium

Table showing the Effects of Temperature on Equilibrium

CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN TEMPERATURE	EQUILIBRIUM MOVES IN THE **ENDOTHERMIC** DIRECTION TO REVERSE THE CHANGE
DECREASE IN TEMPERATURE	EQUILIBRIUM MOVES IN THE **EXOTHERMIC** DIRECTION TO REVERSE THE CHANGE

- **Example:** Iodine monochloride reacts reversibly with chlorine to form iodine trichloride

$$\underset{\text{dark brown}}{ICl} + Cl_2 \rightleftharpoons \underset{\text{yellow}}{ICl_3}$$

- When the equilibrium mixture is heated, it becomes **dark brown** in colour. You can use this observation to deduce whether the backward reaction is exothermic or endothermic
- Equilibrium has shifted to the left as the colour dark brown means that more ICl is produced

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- Increasing temperature moves the equilibrium in the endothermic direction
- **So the backward reaction is endothermic**

Your notes

The Effect of Pressure on Equilibrium

Table showing the Effects of Pressure on Equilibrium

CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN PRESSURE	EQUILIBRIUM SHIFTS IN THE DIRECTION THAT PRODUCES THE **SMALLER** NUMBER OF MOLECULES OF GAS TO DECREASE THE PRESSURE AGAIN
DECREASE IN PRESSURE	EQUILIBRIUM SHIFTS IN THE DIRECTION THAT PRODUCES THE **LARGER** NUMBER OF MOLECULES OF GAS TO INCREASE THE PRESSURE AGAIN

Example: Nitrogen dioxide can form dinitrogen tetroxide, a colourless gas

$$\underset{\text{brown gas}}{2NO_2} \rightleftharpoons \underset{\text{colourless gas}}{N_2O_4}$$

- Predict the effect of an increase in pressure on the position of equilibrium:
 - Number of molecules of gas on the left = 2
 - Number of molecules of gas on the right = 1
- An **increase** in pressure will cause equilibrium to shift in the direction that produces the **smaller** number of molecules of gas
- **So equilibrium shifts to the right**
- The reaction mixture becomes paler as more colourless N_2O_4 is produced

The Effect of Concentration on Equilibrium

Table showing the Effects of Concentration on Equilibrium

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CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN CONCENTRATION	EQUILIBRIUM SHIFTS TO THE **RIGHT** TO REDUCE THE EFFECT OF INCREASE IN THE CONCENTRATION OF A REACTANT
DECREASE IN CONCENTRATION	EQUILIBRIUM SHIFTS TO THE **LEFT** TO REDUCE THE EFFECT OF A DECREASE IN REACTANT (OR AN INCREASE IN THE CONCENTRATION OF PRODUCT)

Your notes

Example: Iodine monochloride reacts reversibly with chlorine to form iodine trichloride

$$\underset{\text{dark brown}}{ICl} + Cl_2 \rightleftharpoons \underset{\text{yellow}}{ICl_3}$$

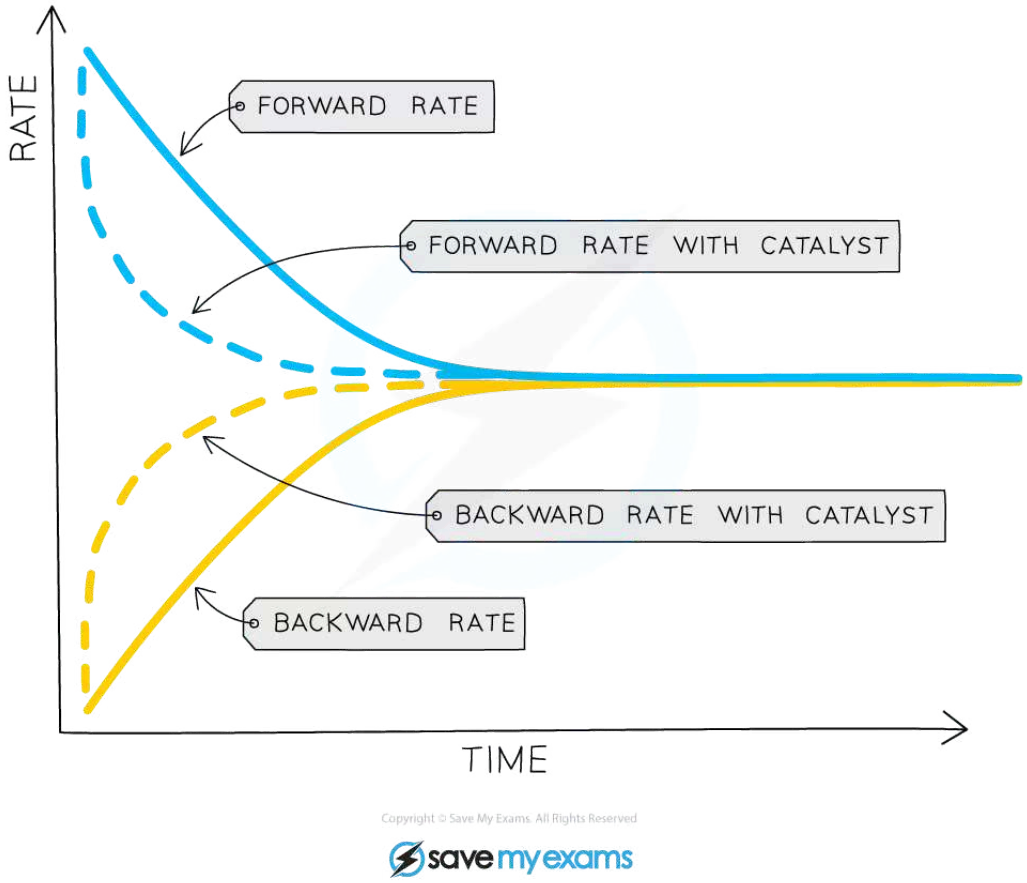
- Predict the effect of an increase in concentration on the position of equilibrium:
 - An increase in the concentration of ICl or Cl_2 causes the equilibrium to shift to the **right** so more of the **yellow** product is formed
 - A decrease in the concentration of ICl or Cl_2 causes the equilibrium to shift to the **left** so more of the **dark brown** reactant is formed

The Effect of a Catalyst on Equilibrium

- The presence of a catalyst does **not** affect the position of equilibrium but it does increase the rate at which equilibrium is reached
- This is because the catalyst increases the rate of **both** the forward and backward reactions by the same amount (by providing an alternative pathway requiring lower activation energy)
- As a result, the **concentration** of reactants and products is nevertheless the **same** at equilibrium as it would be without the catalyst

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Diagram showing the effect of a catalyst on the time taken for equilibrium to be established

Exam Tip

When the conditions at equilibrium are changed, the system always responds by doing the **opposite**. For example if the concentration is increased the system tries to reduce it by changing the direction of the reaction or if the temperature is increased the system will try to reduce the temperature by absorbing the extra heat.

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The Haber Process


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The Haber Process

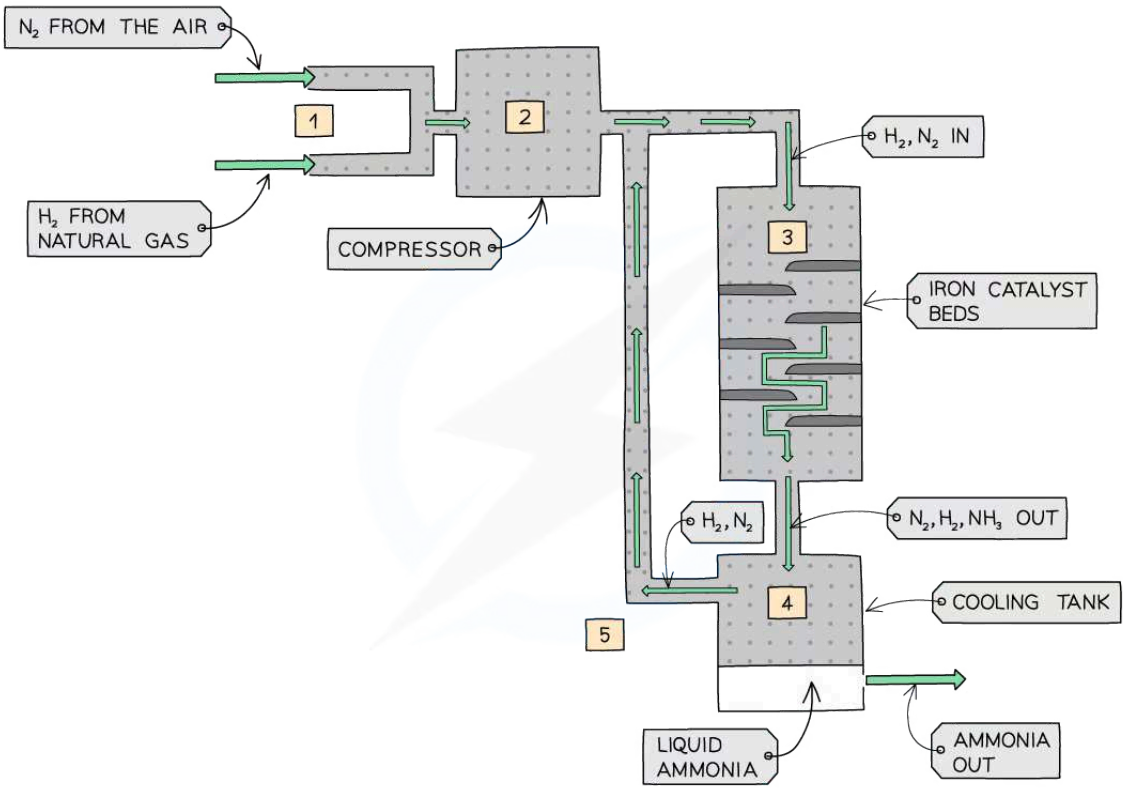
- Ammonia is manufactured in an **exothermic** reaction called the **Haber process** which occurs in five stages:
- **Stage 1:** H_2 and N_2 are obtained from methane in natural gas and the air respectively and are pumped into the compressor through pipe
- **Stage 2:** The gases are compressed to about **200 atmospheres** (20,000 kPa) inside the compressor
- **Stage 3**: The pressurised gases are pumped into a tank containing layers of catalytic **iron** beds at a temperature of **450 °C**. Some of the hydrogen and nitrogen react to form ammonia:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

- **Stage 4:** Unreacted H_2 and N_2 and the product ammonia pass into a cooling tank. The ammonia is liquefied and removed to pressurised storage vessels
- **Stage 5**: The unreacted H_2 and N_2 gases are recycled back into the system and start over again

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The production of ammonia by the Haber process

Exam Tip

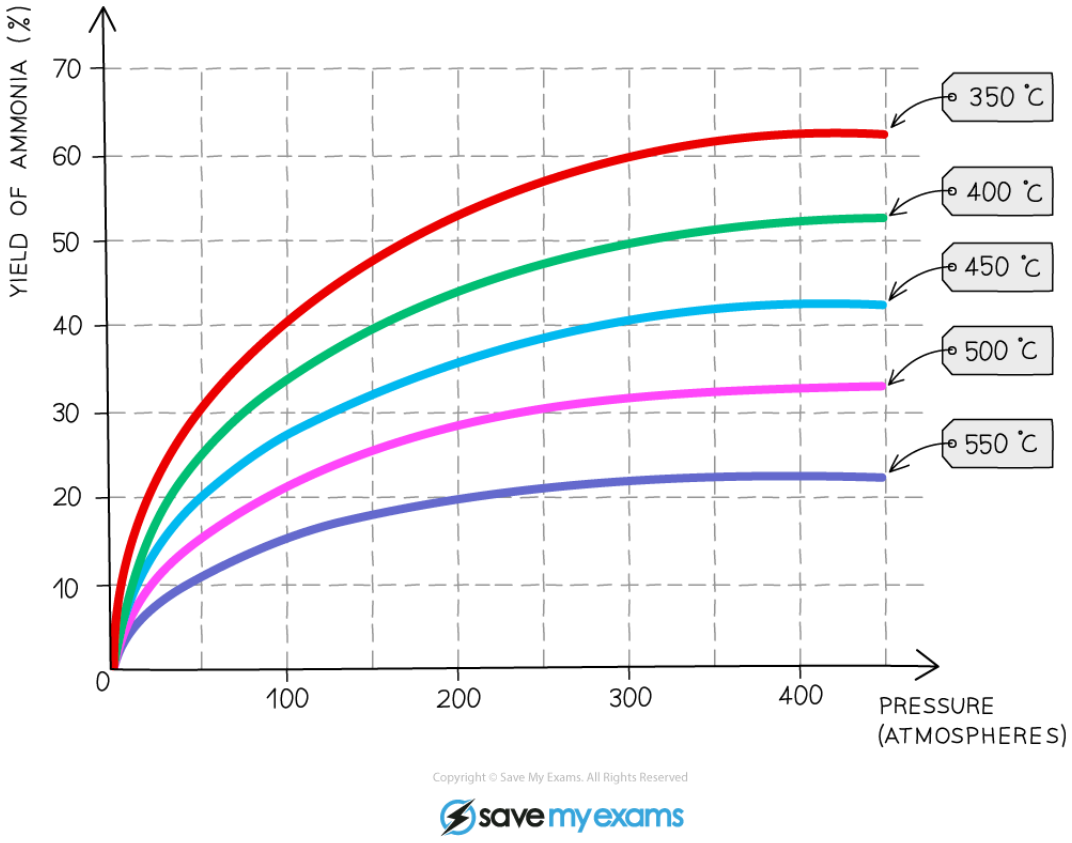
You need to be able to recall the temperature, pressure and name of the catalyst used in the Haber process.

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Explaining the Conditions in the Haber Process

Your notes

- Reaction conditions such as temperature and pressure affect the **rate** of a reaction
- If the reaction is reversible then the **position** of **equilibrium** is also affected by changes in these conditions and often we must consider a trade-off between the rate of reaction and product yield
- The graph below illustrates the effects of changing temperature and pressure on the yield of ammonia obtained
- By following any of the curved lines on the graph it can be seen that as the **pressure increases,** so too does the **yield** at any given temperature
- By following any vertical line upwards from the x-axis, the graph shows that as the temperature **decreases,** the yield actually **increases**
- The actual conditions used must be chosen depending on a number of economical, chemical and practical considerations



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The yield of ammonia produced changes with changes made to temperature and pressure

Economic Considerations

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Your notes

- Like all industries, companies that manufacture and sell chemical goods do so to make a profit
- Part of the industrial process is the economic decision on how and where to design and implement a manufacturing site
- The availability and cost of **raw materials** is a major consideration which must be studied well before any decisions are taken
- In the Haber Process the raw materials are readily available and inexpensive to purify:
 - **Nitrogen** - from the **air**
 - **Hydrogen**- from **methane** in natural gas
- If the cost of extraction of raw materials is too high or they are unavailable then the process is no longer **economically viable**
- Many industrial processes require huge amounts of **heat** and **pressure** which is very expensive to maintain
- Production energy costs are also a factor to be considered carefully and alongside the raw materials issue

Temperature: 450 °C

- A **higher** temperature would favour the reverse reaction as it is endothermic (takes in heat) so a higher yield of **reactants** would be made
- If a **lower** temperature is used it favours the forward reaction as it is exothermic (releases heat) so a higher yield of **products** will be made
- However at a lower temperature the rate of reaction is very **slow**
- So 450 °C is a compromise temperature between having a **lower yield** of products but being made more quickly

Pressure: 200 atm

- A **lower** pressure would favour the reverse reaction as the system will try to increase the pressure by creating more molecules (4 molecules of gaseous reactants) so a higher yield of **reactants** will be made
- A **higher** pressure would favour the forward reaction as it will try to decrease the pressure by creating less molecules (2 molecules of gaseous products) so a higher yield of **products** will be made
- However high pressures can be dangerous and very expensive equipment is needed
- So 200 atm is a **compromise** pressure between a lower yield of products being made **safely** and **economically**

Catalyst: Iron

- The presence of a catalyst does **not** affect the position of equilibrium but it does increase the **rate** at which equilibrium is reached
- This is because the catalyst increases the rate of **both** the forward and backward reactions by the same amount (by providing an alternative pathway requiring **lower activation energy**)
- As a result, the **concentration** of reactants and products is nevertheless the **same** at equilibrium as it would be without the catalyst.

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So a catalyst is used as it helps the reaction reach equilibrium quicker

- It allows for an acceptable yield to be achieved at a **lower temperature** by lowering the activation energy required
- Without it the process would have to be carried out at an even **higher temperature**, increasing **costs** and decreasing **yield** as the higher temperature **decomposes** more of the NH_3 molecules


Your notes

Exam Tip

The reaction conditions chosen for the Haber process are not ideal in terms of the yield but do provide balance between product yield, reaction rate and production cost. These are called **compromise conditions** as they are chosen to give a good compromise between the yield, rate and cost.

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The Contact Process


Your notes

The Contact Process

- Sulfuric acid is synthesised by the **Contact process**
- Concentrated sulfuric acid is used in car batteries, making **fertilisers**, **soaps** and **detergents**
- The first stage is the production of sulfur dioxide, either by burning sulfur to oxidise the sulfur (equation shown below), or roasting sulfide ores

$$S + O_2 \rightarrow SO_2$$

- The **main stage** in the Contact process is the oxidation of sulfur dioxide to sulfur trioxide using a **vanadium(V) oxide**, **V_2O_5, catalyst:**

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

- The oxygen used in this stage is obtained from air
- The conditions for this main stage of production are:
 - A temperature of **450 °C**
 - A pressure of **2 atm (200 kPa)**
- Once sulfur trioxide is formed, it undergoes more processes to produce sulfuric acid

Exam Tip

You need to recall the temperature, pressure and catalyst needed for the Contact process and the equation for the main stage only.

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Explaining the Conditions in the Contact Process


Your notes

- Similar to the Haber process, the pressure and temperature used need to be considered
- The equation for the main stage of the Contact process is:

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

Temperature: 450°C

- The forward reaction is **exothermic**, so increasing the temperature shifts the position of equilibrium to the left in the direction of the **reactants**
- Therefore the higher the temperature, the **lower** the yield of sulfur trioxide
- The optimum temperature is a compromise between a higher rate of reaction at a higher temperature and a lower equilibrium yield at a higher temperature

Pressure: 2 atm

- An increase in pressure shifts the position of equilibrium to the **right** in the direction of a smaller number of gaseous molecules
- However, the position of equilibrium lies far to the right (the equilibrium mixture contains about 96% sulfur trioxide)
- So the reaction is carried out at just above atmospheric pressure because:
 - High pressures can be dangerous and very expensive equipment is needed
 - A higher pressure causes the sulfur dioxide to liquefy

Exam Tip

Remember: These conditions are a **compromise** between yield, rate, safety and cost.

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Cambridge O Level Chemistry

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6.4 Redox

Contents

Oxidation & Reduction

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Oxidation & Reduction


Your notes

Oxidation & Reduction

Redox reactions

- Oxidation and reduction take **place together** at the **same time** in the **same reaction**
- These are called **redox** reactions
- **Oxidation** can be defined as:
 - The gain of oxygen
 - The loss of electrons
 - An increase in oxidation number
- **Reduction** can be defined as:
 - The loss of oxygen
 - The gain of electrons
 - A decrease in oxidation number

Names using oxidation numbers

- Transition elements can bond in different ways by forming ions with different charges
- When naming, the charge on the ion is shown by using a Roman numeral after the element's name
 - e.g. iron can form ions with a 2+ charge, called **iron(II)** ions or a 3+ charge, called **iron(III)** ions
- The Roman numeral is the **oxidation number** of the element
- When iron reacts with oxygen to form iron oxide, the formula depends on the oxidation state of the iron ions
 - The compound where iron has a 2+ charge has the formula FeO and is called iron(II) oxide
 - The compound where iron has a 3+ charge has the formula Fe_2O_3 and is called iron(III) oxide

Exam Tip

You may see the term oxidation state used instead of oxidation number. Although there is a subtle difference between the two terms (this is beyond the scope of this course), they are often used interchangeably. Usually oxidation number is used to refer to the Roman numerals found within the name.

Oxidation and reduction in terms of oxygen

- **Oxidation** is a reaction in which **oxygen** is **added** to an element or a compound
- **Reduction** is a reaction in which **oxygen** is **removed** from an element or compound

Example: Identifying the loss and gain of oxygen in an equation

zinc oxide + carbon → zinc + carbon monoxide

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$$ZnO + C \rightarrow Zn + CO$$

- In this reaction, the zinc oxide has been reduced since it has **lost** oxygen
- The carbon atom has been oxidised since it has **gained** oxygen

Your notes

Oxidation and reduction in terms of electrons and oxidation number

- Redox reactions can also be defined in terms of electron transfer
- **Oxidation** is a reaction in which an element, ion or compound **loses electrons**
 - The **oxidation number** of the element is **increased**
 - This can be shown in a **half equation**, e.g. when silver reacts with chlorine, silver is **oxidised** to silver ions:

$$Ag \rightarrow Ag^+ + e^-$$

- **Reduction** is a reaction in which an element, ion or compound **gains electrons**
 - The **oxidation number** of the element is **decreased**
 - This can be shown in a half equation, e.g. when oxygen reacts with magnesium, oxygen is **reduced** to oxide ions:

$$O_2 + 4e^- \rightarrow 2O^{2-}$$

Example: Identifying Redox Reactions

zinc + copper sulphate → zinc sulphate + copper

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- The ions present (with state symbols) in the equation are:

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

- The **spectator ions** (those that do not change) are $SO_4^{2-}(aq)$
- These can be removed and the ionic equation written as:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

- By analysing the ionic equation, we can split the reaction into two half equations by adding in the electrons to show how the changes in charge have occurred:

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

- It then becomes clear that zinc has been oxidised as it has **lost electrons**
- Copper ions have been reduced as they have **gained electrons**

Exam Tip

Use the mnemonic **OIL-RIG** to remember oxidation and reduction in terms of the movement of electrons: **O**xidation **I**s **L**oss – **R**eduction **I**s **G**ain.

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Identifying Redox Reactions

Oxidation Number


Your notes

- The **oxidation number** (also called oxidation state) is a number assigned to an atom or ion in a compound which indicates the **degree of oxidation** (or reduction)
- It shows the number of electrons that an atom has lost, gained or shared in forming a compound
- The oxidation number helps you to keep track of the movement of electrons in a redox process
- It is written as a +/- sign followed by a number (not to be confused with charge which is written by a number followed by a +/- sign)
- E.g. aluminium in a compound usually has the oxidation state +3
- A few simple rules help guide you through the process of determining the oxidation number of any element

Table of Rules for Assigning Oxidation Numbers

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Your notes

Rule	Example
1. The oxidation number of any uncombined element is zero	H_2 Zn O_2
2. Many atoms or ions have fixed oxidation number in compounds	Group 1 elements are always +1 Group 2 elements are always +2 Fluorine is always −1 Hydrogen is +1 (except for in metal hydrides like NaH, where it is −1) Oxygen is −2 (except in peroxides, where it is −1 and F_2O where it is +2)
3. The oxidation number of an element in mono-atomic ion is always same as the charge	Zn^{2+} oxidation number = +2 Fe^{3+} oxidation number = +3 Cl^- oxidation number = −1
4. The sum of the oxidation number in a compound is zero	NaCl oxidation number of NA = +1 oxidation number of Cl = −1 sum oxidation number = 0
5. The sum of oxidation number in ion is equal to the charge on the ion	SO_4^{2-} oxidation number of S = +6 oxidation number of 4 O atoms = 4 × (−2) sum oxidation number = −2

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- Redox reactions can be identified by the changes in the **oxidation number** when a reactant goes to a product

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Worked example

The equation for the reaction between chlorine and potassium iodide is shown below.

$$Cl_2 + 2KI \rightarrow 2KCl + I_2$$

Identify which species has been:

a) Oxidised

b) Reduced

Answer:

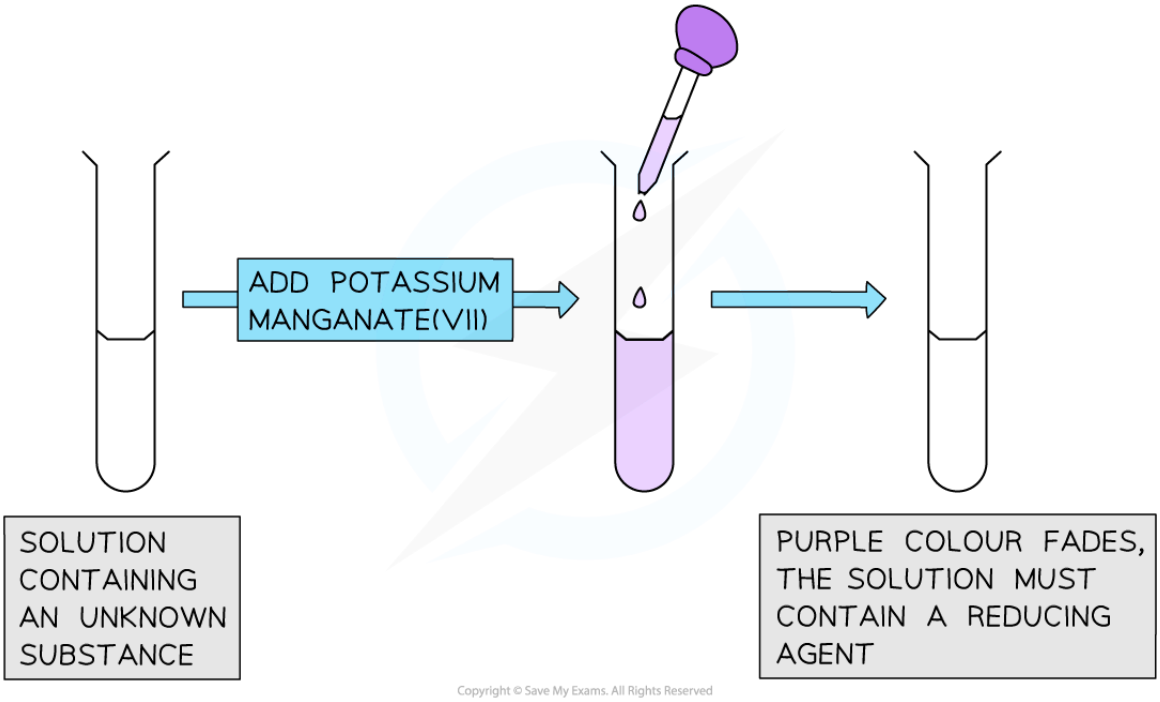
- The species that has been oxidised is iodine
 - The oxidation number of I^- is -1
 - The oxidation number of iodine in I_2 is 0
 - The oxidation number has increased so the iodine has been oxidised (lost electrons)
 - $2I^-(aq) \rightarrow I_2(s) + 2e^-$
- The species that has been reduced is chloride ions
 - The oxidation number of chlorine as Cl_2 is 0.
 - The oxidation number of Cl^- is -1
 - The oxidation number has decreased so the Cl^- has been reduced (gained electrons)
 - $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Identifying Redox Reactions by Colour Changes

- The tests for redox reactions involve the observation of a colour change in the solution being analysed
- Two common examples are acidified potassium manganate(VII), and potassium iodide
- Potassium manganate(VII), $KMnO_4$, is an **oxidising agent** which is often used to test for the presence of **reducing** agents
- When acidified potassium manganate(VII) is added to a reducing agent its colour changes from **purple** to **colourless**

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Diagram to show the colour change when potassium manganate(VII) is added to a reducing agent

- Potassium iodide, KI, is a **reducing** agent which is often used to test for the presence of **oxidising** agents
- When added to an acidified solution of an oxidising agent such as aqueous chlorine or hydrogen peroxide (H_2O_2), the solution turns a **red-brown** colour due to the formation of iodine, I_2:

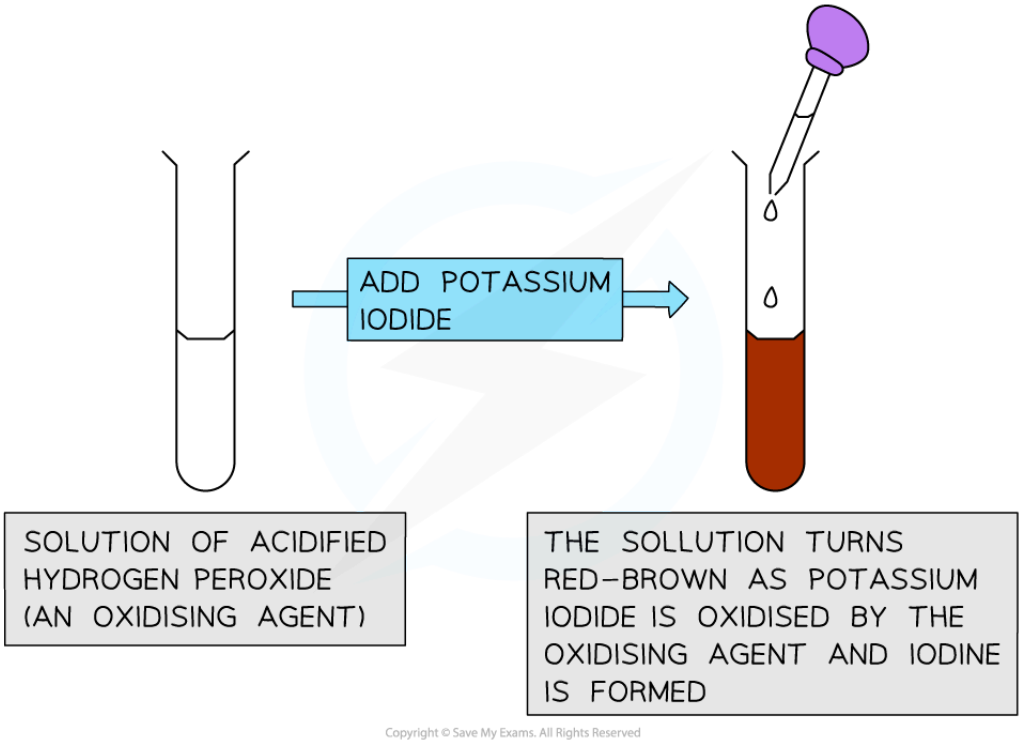
$$2KI\,(aq) + H_2SO_4\,(aq) + H_2O_2\,(aq) \rightarrow I_2\,(aq) + K_2SO_4\,(aq) + 2H_2O\,(l)$$

- The potassium iodide is oxidised as it loses electrons and hydrogen peroxide is reduced, therefore potassium iodide is acting as a **reducing agent** as it will itself be oxidised:

$$2I^- \rightarrow I_2 + 2e^-$$

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Diagram to show the colour change when potassium iodide is added to an oxidising agent

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Oxidising & Reducing Agents


Your notes

Oxidising agent

- A substance that **oxidises** another substance, and becomes **reduced** in the process
- An oxidising agent **gains** electrons as another substance loses electrons
- Common examples include hydrogen peroxide, fluorine and chlorine

Reducing agent

- A substance that **reduces** another substance, and becomes **oxidised** in the process
- A reducing agent **loses** electrons as another substance gains electrons
- Common examples include carbon and hydrogen
- The process of reduction is very important in the chemical industry as a means of extracting metals from their ores

Example

$$CuO + H_2 \rightarrow Cu + H_2O$$

- In the above reaction, hydrogen is reducing the CuO and is itself oxidised as it has lost electrons, so the **reducing agent** is therefore **hydrogen**:

$$H_2 \rightarrow 2H^+ + 2e^-$$

- The CuO is reduced to Cu by gaining electrons and has oxidised the hydrogen, so the **oxidising agent** is therefore **copper oxide**

$$Cu^{2+} + 2e^- \rightarrow Cu$$

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Worked example

When iron reacts with bromine to form iron(II) bromide, a redox reaction reaction occurs:

$$Fe + Br_2 \rightarrow FeBr_2$$

What is acting as the reducing agent in this reaction?

Answer

Step 1 - Write half equations to work out what has gained/lost electrons

$$Fe \rightarrow Fe^{2+} + 2e^-$$

$$Br_2 + 2e^- \rightarrow 2Br^-$$

Fe loses electrons; Br_2 gains electrons

Step 2 - Deduce what has been oxidised/reduced (remember OIL RIG)

Fe has been oxidised as it has lost electrons

Br_2 has been reduced as it has gained electrons

Step 3 - Identify the reducing agent

Fe is the reducing agent as it has been oxidised by losing electrons and caused Br_2 to be reduced as it gained electrons

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Cambridge O Level Chemistry

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7.1 The Characteristic Properties of Acids & Bases

Contents

- Properties of Acids & Bases
- The Ions in Acids & Alkalis
- Proton Transfer, Strong & Weak Acids

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